Subject : Chemistry FREE from www.tekoclasses.com Class : X (CBSE)

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➤➤➤CHEMICAL REACTIONS & CHEMICAL EQUATIONS⮜⮜⮜

1.1 INTRODUCTION:

Chemistry is defined as that branch of science which deals with the composition and properties of matter and the changes that matter undergone by various interactions. A chemical compound is formed as a result of a chemical change and in this process different type of energies such as heat, electrical energy, radiation etc. are either absorbed or evolved. The total mass of the substance remains the same throughout the chemical change.

1.2 CHEMICAL ACTION OR REACTION:

When a chemical change occurs, a chemical action is said to have taken place. A chemical change or chemical action is represented by a chemical equation. The matter undergoing change in known as reactant and new chemical component formed is known as product.

1.2 (a) Characteristics of a Chemical Reaction:

When we heat sugar crystals they melt and on further heating they give steamy vapour, leaving behind brownish black mass. On cooling no sugar crystals appears. Thus change which takes place on heating sugar is a chemical change and the process which brings about this chemical change is called chemical reaction.

- In this reaction the substance which take part in bringing about chemical change are called reactants.
- The substance which are produced as a result of chemical change are called products.
- These reactions involve braking and making of chemical bonds.
- Product(s) of the reaction is/are new substances with new name(s) and chemical formula.
- It is often difficult or impossible to reverse a chemical reaction.
- Properties of products formed during a chemical reaction are different from thos of the reactants.
- Apart from heat other forms of energies are light and electricity which are also used in carrying out chemical changes.

In all chemical reactions, the transformation from reactants to products is accompanied by various characteristics, which are-

(i) Evolution of gas : Some chemical reactions are characterized by evolution of a gas.

- When zinc metal is treated with dilute sulphuric acid, hydrogen gas is evolved. The hydrogen gas burns with a pop sound.
  $Zn\ (s) + H_2SO_4\ (dilute) \rightarrow ZnSO_4\ (aq) + H_2(g)$

- When washing soda is treated with hydrochloric acid, it gives off colorless gas with lots of effervescence.
  $Na_2CO_3(s) + 2HCl \rightarrow 2NaCl\ (aq) + H_2O(l) + CO2(g)$

- $2NaNCO_3\ (s) \xrightarrow{heat} Na_2SO_3\ (s) + H_2O\ (\ell) + CO_2\ (g)$
  Sodium hydrogen carbonate → Sodium carbonate + Water + Carbon dioxide

(ii) Change of colour: Certain chemical reactions are characterized by the change in colour of reacting substance.

- When red lead oxide is heated strongly it forms yellow coloured lead monoxide and gives off oxygen gas.
  $2Pb_3O_4\ (s) \xrightarrow{heat} 6PbO(s) + O_2(g)$
  Lead oxide (Red) → Lead monoxide (Yellow)

- When copper carbonate (green) is heated strongly it leaves behind a black residue.
  $CuCO_3\ (s) \xrightarrow{heat} CuO(s) + CO_2\ (g)$
  Copper carbonate (Green) → Copper oxide (Black) + Carbon dioxide

- $2Pb(NO_3)_2(s) \xrightarrow{heat} 2\ PbO(s) + 4NO_2\ (g) + O_2\ (g)$
  Lead (II) nitrate (White) → Lead (II) oxide (Yellow) + Nitrogen dioxide (Brown)

- $C_{12}H_{22}O_{11}\ (s) \xrightarrow{heat} 12C(s) + 11H_2O$
  White sugar → Carbon Black + Water

(iii) Formation of precipitate : Some chemical reactions are characterized by the formation of precipitate (an insoluble substance), when the solutions of the soluble chemical compounds are mixed together.

- When silver nitrate solution is mixed with a solution of sodium chloride.

  $AgNO_3\ (aq) + NaCl\ (aq) \longrightarrow NaNO_3\ (aq) + AgCl\ (s)$
  Silver nitrate (Colourless) + Sodium chloride (Colourless) → Sodium nitrate (Colourless) + Silver chloride (White precipitate)

- A dirty green precipitate of ferrous hydroxide is formed, when a solution of ferrous sulphate is mixed with sodium hydroxide solution.

$FeSO_4$ (aq) +	$2NaOH(aq)$ $\longrightarrow$	Na_2SO_4 (aq) +	$Fe(OH)_2$ (aq)
Ferrous sulphate (Light green)	sodium hydroxide (Colourless)	Sodium sulpahte (Colourless)	Ferrous hydroxide (Dirty green precipitate)

- $BaCl_2$ (aq) + dill H_2SO_4 $\longrightarrow$ $BaSO_4$ (s) + 2HCl (aq)

Bacl$_2$ (aq) +	dill H_2SO_4 $\longrightarrow$	$BaSO_4$ (s) + 2HCl (aq)
Barium chloride		Barium sulphate (White precipitate)

(iv) Energy changes : all chemical reactions proceed either with the absorption or release of energy. One the basis of energy changes, there are two types of reactions:

(A) Endothermic reaction : A chemical reaction which is accompanied by the absorption of heat energy is called an endothermic reaction.

- $C\ (s) + 2S\ (s) \xrightarrow{\text{Heat}} CS_2\ (\ell)$

- Light energy is essential for biochemical reaction, photosynthesis, by which green plants prepare their food from carbon dioxide & water.

(B) Exothermic reaction : A chemical reaction which is accompanied by the release of heat energy is called exothermic reaction.
When magnesium wire is heated from its tip in a bunsen flame, it catches fire and burns with a dazzling white flame with release of heat and light energy.

$2Mg\ (s) + O_2\ (g) \xrightarrow{\text{Heat}} 2MgO\ (s) + \text{Energy}$

- When quick lie (calcium oxide) is placed in water, the water becomes very hot and sometimes starts boiling. It is because of release of heat energy during the reaction.

CaO (s) +	H_2O $\longrightarrow$	$Ca(OH)_2$ (aq) + Heat energy
Calcium oxide	Water	Calcium hydroxide

(v) Change of state: Some chemical reactions are characterised by a change in state i.e. solid, liquid or gas

- Two volumes of hydrogen gas react with one volume of oxygen gas to from water.

$2H_2\ (g) + O_2\ (g) \rightarrow 2H_2O\ (\ell)$

or when electric current is passed through water it splits into its elements.

$2H_2O\ (\ell) \xrightarrow{\text{Electric current}} 2H_2(g) + O_2\ (g)$

NH_3 (g) +	HCl (g) $\longrightarrow$	NH_4Cl (s)
Ammonia	Hydrochloric acid	Ammonium Chloride

1.3 CHEMICAL EQUATIONS :

All chemical changes are accompanied by chemical reactions. These reactions can be described in sentence form, but the description would be quite long. Chemical equations have been framed to describe the chemical reactions.

A chemical equation links together the substance which react (reactants) with the new substances that are formed (products).

Zinc + Hydrochloric acid $\longrightarrow$ Zinc chloride + Hydrogen
(Reactants) (Products)

A Chemical reaction can be summarised by chemical equation.

1.3 (a) Types of Chemical Equations :

(i) Word equations : A word equation links together the names of the reactants with those of the products. For example, the word equation, when magnesium ribbon burns in oxygen to form a white powder of magnesium oxide, may be written as follows-

Magnesium + Oxygen $\longrightarrow$ Magnesium oxide
(Reactants) (Product)

Similarly, the word equation for the chemical reaction between granulated zinc and hydrochloric acid may be written as -
Zinc + Sulphuric acid $\rightarrow$ Zinc sulphate + Hydrogen

In a word equation

- The reactants are written on the left hand side with a plus sign (+) between them.
- The products are written on the right hand side with a plus sign (+) between them.
- An arrow ($\rightarrow$) separates the reactants from the products.
- The direction of the arrow head points towards the product.

Although word equations are quite useful, yet they don't give the true picture of the chemical reactions.

(ii) Symbol equation : A brief representation of a chemical reaction in terms of symbols and formulae of the substance involved is known as a symbol equation.
In a symbol equation, the symbols and formulae of the elements and compounds are written instead of their word names.

For e.g. Burning of magnesium in oxygen to form magnesium oxide may be written as follows :
$Mg + O_2 \longrightarrow MgO$

Symbol equations are always written from the word equations.

1.3 (b) Unbalanced and Balanced Chemical Equations :

In an unbalanced equation, the number of atoms of different elements on both side of the equation are not equal. For example, in the equation given below, the number of Mg atoms on both sides of the equation is one (same), but the number of oxygen atoms are not equal, It is known as an unbalanced equations.

$$Mg + O_2 \longrightarrow MgO$$

An unbalanced equation is also called skeletal equation.

In a balanced equating, the number of different elements on both sides of the equation are always equal. The balanced equation for the burning of magnesium ribbon in oxygen is written as -

$2\ Mg + O_2 \longrightarrow 2\ MgO$

(i) Importance of balanced chemical equation: The balancing of a chemical equation is essential or necessary to fulfill the requirement of "Law of conservation of mass".

(ii) Balancing of chemical equations: Balancing of chemical equations may be defined as the process of making the number of different types of elements, on both side of the equations, equal.

The balancing of a chemical equation is done with the help of **Hit and Trial method. In this method,** the coefficients before the symbols or formulae of the reactants and products are adjusted in such a way that the total number of atoms of each element on both the side of the arrow head become equal. This balancing is also known as mass balancing because the atoms of elements on both side are equal and their masses will also be equal.

The major steps involved in balancing a chemical equation are as follow –

- Write the chemical equations in the form a word equations. Keep the reactants on the left side and the products on the right side. Separate them by an arrow whose head ($\rightarrow$) points from the reactants towards the product.

- Convert the word equation into the symbol equation by writing the symbols and formulae of all the reactants and product.

- Make the atoms of different elements on both side of the equation equal by suitable method. This is known as balancing of equation.

- Do not change the formulae of the substance while balancing the equation.

- Make the equations more informative if possible.

Example :

1. Zinc reacts with dilute sulphuric acid to give zinc sulphate and hydrogen.

Solution : The word equation for the reaction is -

Zinc + Sulphuric acid $\rightarrow$ Zinc sulphate + Hydrogen

The symbol equation for the same reactions is -

$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

Let us count the number of atoms of all the elements in the reactants and products on both sides for the equations.

Element	No. of atoms of reactants (L.H.S.)	No. of atoms of products (R.H.S.)
Zn	1	1
H	2	2
S	1	1
O	4	4

As the number of atoms of the elements involved in the reactants and products are equal, the equation is already balanced.

2. Iron reacts with water (steam) to form iron (II, III) oxide and liberates hydrogen gas.

Solution :- The word equation for the reactions is -

Iron + Water $\rightarrow$ iron (II, III) oxide + Hydrogen

The symbol equation for the same reaction is-

$Fe + H_2O \rightarrow Fe_3O_4 + H_2$

The balancing of the equations is done is the following steps:

I : Let us count the number of atoms of all the elements in the reactants and products on both sides of the equation.

Element	No. of atoms of reactants (L.H.S.)	No. of atoms of products (R.H.S.)
Fe	1	3
H	2	3
O	2	4

Thus, the number of H atoms are equal on both sides, At the same time, the number of Fe and O atoms are not equal.

II : On inspection, the number of O atoms in the reactant (H_2O) is 1 while in the product (Fe_3O_4), these are 4. To balance the atoms, put coefficient 4 before H_2O on the reactant side. The partially balance equation may be written as

$Fe + 4H_2O \rightarrow Fe_3O_4 + H_2$

III : In order to equate H atoms, put coefficient 4 before H_2 on the product side, As a result, the H atoms on both side on of the equation become 8 and are thus balanced. The partially balanced equation may now be written as

$Fe + 4H_2O \rightarrow Fe_3O_4 + H_2$

IV : In order to balance the Fe atoms, put coefficient 3 before Fe on the reactant side. The equation formed may be written as -

$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

V : on final inspection, the number of atoms of all the elements on both sides of the equation are equal. Therefore, the equation is balanced.

1.3 (c) Writing State Symbols:

The chemical equations or symbol equations which we have enlisted don't mention the physical states of the reactant and product species involved in the reaction. In order to make the equation more informative, the physical state are also mentioned with the help of certain specific symbols known as state symbols. These symbols are

- (s) for solid state
- (ℓ) for liquid state
- (g) for gaseous state
- (aq) for aqueous solution i.e., solution prepared in water.

Sometimes a gas if evolved in a reaction is shown by the symbol ($\uparrow$) i.e., by an arrow pointing upwards. Similarly the precipitate, if formed during the reaction, is indicated by the symbol ($\downarrow$) i.e., by an arrow pointing downwards.
The abbreviation 'ppt' is also use to represent the precipitate, if formed.

(i) $2Na(s) + 2H_2O(\ell) \rightarrow 2NaOH\ (aq) + H_2(g)$ or $H_2(\uparrow)$

(ii) $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(\downarrow) + H_2O(\ell)$

(iii) $AnNo_3(aq) + NaCl(aq) \rightarrow AgCl(\downarrow) + NaNO_3\ (aq)$

1.3 (d) Significance of State Symbols:

The state symbols are of most significance for those chemical reactions which are either accompanied by the evolution of heat (exothermic) or by the absorption of heat (endothermic). For example.

$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell) + 572\ kJ$

$2H_2\ (g) + O_2(g) \rightarrow 2H_2O(g) + 44\ kJ$

Both these reactions are of exothermic nature because heat has been evolved in these. Howeve, actual amounts of heat are different when water is in the liquid state i.e. $H_2O(\ell)$ and when it is in the vapour state.

1.3 (e) Specialties of Chemical Equation :

(i) We get the information about the substance which are taking part and formed in the reaction.

(ii) We get the information about the number of molecules of elements or compounds which are either taking part or formed in the chemical reaction.

(iii) We also get the information of weight of reactant or products.

For example - $CaCO_3 \longrightarrow CaO + CO_2$
(100gm) (56 gm) (44 gm)

Total weight of reactants is equal to the total weight of products because matter is never destroyed. In the above example total weight of calcium carbonate (reactant) is 100 gram and of product is also 100 g (56 gram + 44 gram).

(iv) In a chemical equation if any reactant or product is in gaseous state, then its volume can also be determined. For example in the above reaction volume of carbon dioxide is 22.4 liters.

(vi) In a chemical equation with the help of product we can get information about the valency as well.

For example

$Mg + 2HCl \rightarrow MgCl_2 + H_2 (\uparrow)$

In the above reaction one atom of Mg displaces two atoms of hydrogen, so valency of magnesium is two.

All chemical equations are written under N.T.P. Conditions (at 273 K and 1 atmosphere pressure) if conditions are not otherwise mentioned.

1.3 (f) Limitations of Chemical Equations :

(i) We do not get information about the physical state of reactants and products.
For example solid, liquid or gas.

(ii) No information about the concentration of reactants and products is obtained.

(iii) No information about the speed of reaction and sense of timing can be obtained.

(iv) Information regarding the favorable conditions of the reactions such as pressure, temperature, catalyst etc. can't be obtained during the reaction.

(v) We do not get information whether heat is absorbed or evolved during the reaction.

(vi) We do not get information whether the reaction of reversible or irreversible.

(vii) We do not get information about the necessary precautions to be taken for the completion of reaction.

The above limitations are rectified in the following manner –

- The physical sate of reactants and products are represented by writing them in bracket.
- The precipitate formed in the reaction is represented by $(\downarrow)$ symbol and gaseous substance by $(\uparrow)$ symbol.
- To express the concentration, dilute or conc. is written below the symbol.

$$Mg + \underset{\text{(dilute)}}{H_2SO_4} \longrightarrow MgSO_4 + H_2$$

- Favorable conditions required for the completion of reaction are written above and below the arrow.

$$N_2 + 3H_2 \xrightarrow[200\ \text{atm}]{500^0.\text{Fe / Mo}} 2NH_3 + 22400 \text{ Calorie heat.}$$

- Reversible reaction is represented by $(\rightleftharpoons)$ symbol and irreversible reaction by $(\rightarrow)$ symbol.
- The heat absorbed in the chemical reaction is written on the right side by putting negative (-) sign and heat evolved in the chemical reaction is written on the right side by putting positive (+) sign.

$N_2 + 3H_2 \rightleftharpoons 2NH_3 + 22400$ Calorie (Exothermic Reaction)

$N_2 + O_2 \rightleftharpoons 2NO - 43200$ Calorie (Endothermic Reaction)

DAILY PRACTIVE PROBLMES # 1

OBJECTIVE DPP-1.1

1. In the balanced equation -
$aFe_2O_3 + bH_2 \longrightarrow cFe + dH_2O$
The value of a,b,c,d are respectively -
(A) 1,1,2,3 (B) 1,1,1,1 (C) 1,3,2,3 (D) 1,2,2,3

2. Which of the following reactions is not balnced \
(A) $2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$ (B) $2C_4H_{10} + 120_2 \longrightarrow 8CO_2 + 10H_2O$
(C) $2Al + 6H_2O \longrightarrow 2Al\ (OH)_3 + 3H_2$ (D) $4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$

3. The equation - $Cu + xHNO_3 \rightarrow Cu(NO_3)_2 + yNO_2 + 2H_2O$
The values of x and y are-
(A) 3 and 5 (B) 8 and 6 (C) 4 and 2 (D) 7 and 1

4. Neutralization reaction is an example of -
(A) exothermic reaction (B) endothermic reaction
(C) oxidation (D) none of these

5. Which of the following statements is/are true \
(A) The total mass of the substance remains same in a chemical change.
(B) A chemical change is permanent and irreversible.
(C) A physical change is temporary and reversible.
(D) All the these.

6. Which of the following statements is correct
(A) A chemical equation tells us about the substances involved in a reaction.
(B) A chemical equation informs us about the symbols and formulae of the substances involved in a reactin.
(C) A chemical equation tells us about the atoms or molecules of the reactants and products involved in a reaction.
(D) All are correct.

7. $Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$ is an example of-
(A) precipitation reaction (B) endothermic reaction
(C) evolution of gas (D) change in colour

8. When dilute hydrochloric acid is added to iron fillings -
(A) hydrogen gas and ferric chloride are produced.
(B) chlorine gas and ferric hydroxide are produced.
(C) no reaction takes place.
(D) iron salt and water are produced.

9. In the reaction $xPb\ (NO_3)32 \xrightarrow{Heat} yPbo + zNO_2 + O_2$ x,y and z are -
(a) 1,1,2 (B) 2,2,4 (C) 1,2,4 (D) 4,2,2

10. In the reaction $FeSo_4 + x \longrightarrow Na_2SO_4 + Fe(OH)_2$, x is -
(A) Na_2SO_4 (B) H_2SO_4 (C) NaOH (D) None of these

SUBJECTIVE DPP-1.2

1. Balance the following equations -

   (i) $HgO \longrightarrow Hg + O_2$

   (ii) $C_4H_{10}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(\ell)$

2. What are chemical equations? Give significance and limitations of chemical equations ?

3. What information do we get from a chemical equation ? Explain with the help of examples.

4. Write the balanced chemical equations for the following chemical reactions -
   (i) Aqueous solution of sulphuric acid and sodium hydroxide reacts to from aqueous sodium sulphate an water.
   (ii) Phosphorus burns in chlorine gas to from phosphorus pentachloride.

5. Write the balance chemical equations for the following reactions -

   (i) Zinc carbonate (s) $\longrightarrow$ Zinc oxide (s) + Carbon dioxide (g)

   (ii) Potassium bromide (aq) + Barium iodide (aq) $\longrightarrow$ Potassium iodide (aq) + Barium bromide (aq)

6. What happens when electric current is passed through slightly acidic water ?

7. What happens when silver nitrate is mixed with a solution of sodium chloride ?

8. What do you mean by exothermic reactions ? Explain with an example.

9. What do you mean by endothermic reactions ? Explain with an example .

➤➤➤CHEMICAL REACTIONS & CHEMICAL EQUATIONS⮜⮜⮜

2.1 TYPES OF CHEMICAL REACTIONS:

2.1 (a) Addition Reactions :

It is a union of two or more than two substances to from a new substance. It may be brought about by the application of heat, light electricity or pressure.

For eg. $H_2 + Cl_2 \rightarrow 2HCl$

In the above example H_2 and Cl_2 two elements combine to from hydrogen chloride.

Addition reactions may be formed in the following conditions -

(i) When two or more elements combine to form a new compound.

Synthesis reaction : It is a type of addition reaction in which a new substance is formed by the union of its component elements.

For eg. $N_2 + 3H_3 \rightarrow 2NH_3$ (Haber's Process)

Ammonia is synthesised from its components, nitrogen and hydrogen, so it is a synthetic reaction.

All synthesis reaction are addition reactions but all addition reactions are not synthesis reactions.

Other Example of synthesis reactions are -

- $2H_2 + O_2 \longrightarrow 2H_2O$
- $2Mg + O_2 \longrightarrow 2MgO$
- $2Na + Cl_2 \longrightarrow 2NaCl$

(ii) When two or more compounds combine to from a new compound.

For eg.

- $NH_3 + HCl \longrightarrow NH_4Cl$
- $CaO + CO_2 \longrightarrow CaCO_3$
- $CH_2 = CH_2 + Br_2 \longrightarrow \begin{array}{l} CH_2 - Br \\ | \\ CH_2 - Br \end{array}$

(iii) When and element and a compound combine to from a new compound.

For eg.

- $2CO + O_2 \longrightarrow 2CO_2$
- $2CO_2 + O_2 \longrightarrow 2CO_3$

Only single substance is formed as a product in the addition reactions.

2.1 (b) Decomposition Reaction :

It is breaking up of a substance into simpler compounds and it may be brought about by the application of heat, light, electricity etc.

(i) A decomposition reaction brought by heat is known as thermal decomposition.
For eg.

- $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$
- $2Pb(NO_3)_2 \xrightarrow{\Delta} 2PbO + 4NO_2 + O_2$

(ii) Decomposition performed by electricity is known as electrolysis.
For eg.

- $2H_2O \xrightarrow{\text{Electricity}} 2H_2 + O_2$
- $2NaCl \xrightarrow{\text{Electricity}} 2Na + Cl_2$
- $2Al_2O_3 \xrightarrow{\text{Electricity}} 4\,Al + 3O_2$

(iii) A decomposition reaction brought by light is known as photo decomposition.
For eg.

- $2AgBr \xrightarrow{\text{Light}} 2Ag + Br_2$
- $2AgCl \xrightarrow{\text{Light}} 2Ag + Cl_2$

(iv) Decomposition reaction in which a compound decomposes into its elements is known as analysis reaction.
For eg.

- $2HgO \xrightarrow{\Delta} 2Hg + O_2$
- $2HI \xrightarrow{\Delta} H_2 \uparrow + \downarrow$

All analysis reactions are decomposition reactions, but all decomposition reactions are not analysis reactions.

Decomposition reaction is just opposite of the addition reaction.

2.1 (c) Displacement Reactions :

It involves displacement of one of the constituents of a compound by another substance and may be regarded as a displacement reaction.

For eg.

(i) Zinc displaces hydrogen from sulphuric acid.
$Zn\ (s) + \text{dill. } H_2SO_4\ (aq) \longrightarrow ZnSO_4\ (aq) + H_2 \uparrow$

(ii) Iron displaces copper from a copper sulphate solution.
$Fe\ (s) + CuSO_4(aq) \longrightarrow FeSO_4\ (aq) + Cu$

Store in your memory

In general a more reactive element displaces a less reactive element from the soluble solution of its salt.

2.1 (d) Double Displacement :

It is mutual exchange of the radicals of two compounds taking part in the reaction and results in the formation of two new compounds.

- $NaCl\ (aq) + AgNO_3\ (aq) \longrightarrow AgCl \downarrow + NaNO_3\ (aq)$
- $BaCl_2\ (aq) + Na_2SO_4\ (aq) \longrightarrow BaSO_4 \downarrow + 2NaCl\ (aq)$


Store in your memory

Acid base neutralisation reactions are double displacement reactions.

DAILY PRACTICE PROBLEMS # 2

OBJECTIVE DPP-2.1

1. Chemical reaction $2Na + Cl_2 \longrightarrow 2\ NaCl$ is an example of -
(A) Combination reaction
(B) decomposition reaction
(C) displacement reaction
(D) double displacement reaction

2. Which of the following equations is representing combination of two elements?
(A) $CaO + CO_2 \longrightarrow CaCO_3$
(B) $4\ Na + O_2 \longrightarrow 2Na_2O$
(C) $SO_2 + 1/2\ O_2 \longrightarrow SO_3$
(D) $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

3. Which of the following equations is not an example of single displacement reaction?
(A) $2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 23Fe$
(B) $Ca + CO_2 \longrightarrow CaCl_2$
(C) $2KI + Cl_2 \longrightarrow 2KCl + I_2$
(D) $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

4. Which of the following is/are a decomposition reaction(s)?
(A) $2HgO \xrightarrow{Heat} 2Hg + O_2$
(B) $CaCO_3 \xrightarrow{Heat} CaO + CO_2$
(C) $2H_2O \xrightarrow{Electrolysis} H_2 + O_2$
(D) All of these

5. Match the following -

Column A **Types of chemical reaction**	**Column B** **Chemical equations**
(a) Combination reaction	(i) $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$
(b) Decomposition reaction	(ii) $2H_2O \xrightarrow{Electricity} 2H_2 + O_2$
(c) Displacement reaction	(iii) $CaO + CO_2 \longrightarrow CaCO_3$
(d) Analysis reaction	(iv) $Fe + CuSO_4\ (aq.) \longrightarrow FeSo_4(aq) + Cu$

(A) a(ii), B(i), C9iv), d(iii)
(B) a(i), b(ii), c(iii), d(iv)
(C) a(iii), b(i), c(iv), d(ii)
(D) a(iii), b(i), c(iii), d(iv)

6. Which of the following reactions is/are a double displacement reactions (s) ?

(i) $AgNO_3 + NaBr \longrightarrow NaNO_3 + AgBr$

(ii) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$

(iii) $As_4O_4 + 3H_2S \longrightarrow As_2S_3 + 3H_2O$

(iv) $NaOH + HCl \longrightarrow NaCl + H_2O$

(A) (i) & (ii) (B) only (iii) (C) only (iv) (D) (i) to (iv) all

7. $AgNo_3$ (a) + NACl (Aq) $\longrightarrow$ AgCl (s) + $NaNO_3$ (Aq)
Above reaction is a -
(A) precipitation reaction
(B) dboule displacement reaction
(C) combination reaction
(D) (A) and (B) both

8. $H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$
Above equation is a
(i) neutralization reaction
(ii) double displacement reaction)
(iii) decomposition reaction
(iv) addition reaction
(A) (i) to (iv) all
(B) (i) and (ii)
(C) (i) and (iii)
(D) (ii) and (iv)

9. $Zn + H_2SO_4$ (dil) $\longrightarrow ZnSO_4 + H_2\uparrow$
Above equation is a=
(A) Decomposition
(B) Single displacement reaction
(C) Combination reaction
(D) Synthesis reaction

10. The reaction in which two compounds exchange their ions to form two new compounds is-
(A) a displacement reaction
(B) a decomposition reaction
(C) an addition reaction
(D) a double displacement reaction

SUBJECTIVE DPP-2.2

1. Classify the following reactions -

(i) $N_2 + O_2 \longrightarrow 2NO$ - Heat

(ii) $2HgO \longrightarrow 2Hg + O_2$

(iii) $Na_2SO_4 + BaCl_2 \longrightarrow 2NaCl + BaSO_4$

(iv) $CuSO_4$ (aq.) + Zn $\longrightarrow$ $ZnSO_4$ (aq.) + Cu

(v) $NH_3 + HCl \longrightarrow NH_4Cl$

2. Differentiate between combination and synthesis reaction with example.

3. What is an analysis reaction? Give an example.

4. When a white compound 'X' is placed under sunlight, it turns grey, Give the name of reaction and write the balanced chemical equation.

5. What is the difference between displacement and double displacement reaction ? Write equations for these reactions.

6. What happens when copper metal is dipped in silver nitrate solution ? Give the balanced chemical equation for the change.

7. What happens when ferrous sulphate is heated ? Write the name and balanced chemical equation for the change.

8. What happens when the iron nail is kept into copper sulphate solution ?

CHEMICAL REACTIONS & CHEMICAL EQUATIONS

3.1 OXIDATION AND REDUCTION :

3.1 (a) Oxidation :

Oxidation is a chemical reaction in which a substance gains oxygen or loses hydrogen. Since oxygen is an electronegative element and hydrogen is an electropositive element, so, oxidation is defined as a reaction in which a substance gains and electronegative radical or loses and electropositive radical.

(i) A reaction in which a substance gains oxygen is known as oxidation.
For eg.

- $S + O_2 \longrightarrow SO_4$
- $2SO_2 + O_2 \longrightarrow 2SO_3$
- $2Ca + O_2 \longrightarrow 2CaO$
- $Pbs + 2O_2 \longrightarrow PbSO_4$

(ii) Gain or addition of a electronegative radical
For eg.

- $2FeCl_2 + Cl_2 \longrightarrow 2FeCl_3$
- $Mg + Cl_2 \longrightarrow MgCl_2$
- $2FeSO_4 + H_2SO_4 + [O] \longrightarrow Fe_2(SO_4)_3 + H_2O$
- $SnCl_2 + Cl_2 \longrightarrow SnCl_4$

(iii) Removal of a hydrogen atom.
For eg.

- $2HCl \longrightarrow Cl_2 + H_2$
- $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$

(iv) Removal or loss of electropositive radical or element.
For e.g.

- $2KI + H_2O_2 \longrightarrow 2KOH + I_2$

3.1 (b) Reduction :

It is a chemical reaction in which there is a gain of hydrogen or any electropositive radical or a loss of oxygen or electronegative radical.

(i) Gain of hydrogen.

For eg.

- $Cl_2 + H_2S \longrightarrow 2HCl + S$
- $O_2 + 2H_2 \longrightarrow 2H_2O$
- $C_2H_4 + H_2 \longrightarrow C_2H_6$

(ii) Gain of any electropositive radical or element.
For eg.

- $SnCl_2 + 2HgCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$
- $CuCl_2 + Cu \longrightarrow Cu_2Cl_2$

(iii) Loss of oxygen atom.
For eg.

- $CuO + H_2 \longrightarrow Cu + H_2O$
- $ZnO + C \longrightarrow Zn + CO$

(iv) Loss of electronegative radical.
For eg.

- $Fe_2(SO_4)_3 + H_2 \longrightarrow 2FeSO_4 + H_2SO_4$
- $SnCl_4 + Hg_2Cl_2 \longrightarrow 2HgCl_2 + SnCl_2$

3.2 REDOX REACTIONS :

Reduction is loss of electronegative element or radical. From all above example it is clear that oxidation and reduction occur side by side, i.e. there can be no oxidation without and equivalent reduction. In a reaction whenever one substance is oxidised the other is definitely reduced. The reverse is also true whenever one substance is reduced the other is oxidized. Such reactions in which oxidation and reduction take place simultaneously are known as **redox reactions.**

Reduction

$CuO + H_2 \longrightarrow Cu + H_2O$

Oxidation

When hydrogen gas is passed through not cupric oxide, hydrogen is oxidised to water (H_2O) while cupric oxide is reduced to metallic copper by loss of oxygen. Hydrogen gas helps in reduction of cupric oxide to metallic copper so it is known as reducing agent, where as cupric oxide helps in oxidation of hydrogen so it is known as oxidizing agent. A substance, which brings about reduction, is called reducing agent. A substance, which brings about oxidation, is called an oxidizing agent.

3.2 (a) Electronic Interpretation of Oxidation:

The electronic theory attempts to interpret oxidation on the basis of electron transfer. According to octet rule, atom will try to complete its octet by losing gaining or sharing electrons. Sodium chloride is an electrovalent compound and consists of an ion pair (Na^+) (Cl^-) even in the solid state. In its formation, the neutral sodium loses and electron and becomes positively charged sodium ion. Sodium is said to be oxidised and loss of electrons is termed as oxidation.

$2Na \rightarrow 2Na^+ + 2e^-$

$2Na^+ + 2Cl^- \rightarrow 2NaCl$

3.2 (b) Electronic Interpretation of Reduction :

Reduction which is also referred to as electronation is a process involving the gain of electrons and is the reverse of oxidation.

For example

Mg combines with oxygen and is oxidized to MgO. According to electronic theory magnesium atom loses two electrons from its outermost shell (M) and is oxidised to mG which oxygen atom gains these two electrons and gets reduced to oxide anion, hence oxidation involves loss of electrons and it is also referred as de- electronation. Reduction involves gain of electrons so it is referred to as electronation.

$2Mg + O_2 \rightarrow 2MgO$

$Mg \rightarrow Mg^{+2} + 2e^-$

$O + 2e^- \rightarrow O^{2-}$

$Mg^{+2} + O^{2-} \rightarrow MgO$

3.3 EFFECT OF OXIDATION REACTIONS IN EVERYDAY LIFE :

We are all aware of the fact that oxygen is most essential for sustaining life. One can live without food or even water for a number of days but not without oxygen. It is involved in a variety of actions which have wide range of effects on our daily life. Most of them are quite useful while a few may be harmful in nature. Some of these effects are briefly discussed. Some examples are-

3.3 (a) Combustion Reactions:

A chemical reaction in which a substance burns or gets oxidised in the presence of air or oxygen in called combustion reaction. For example, kerosene, coal, charcoal, wood etc. burn in air and thus, undergo combustion. Methane (CH_4) a major constituent of natural gas undergoes combustion in excess of oxygen upon heating.

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$

Methane

Similarly, butane (C_4H_{10}) the main constituent of L.P.G. also undergoes combustion.

$C_4H_{10}(g) + 13/2O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

Butane

All combustion reactions are of exothermic nature and are accompanied by release of heat energy. The human body may be regarded as a furnace or machine in which various food stuffs that we eat undergo combustion or oxidation. The heat energy evolved keeps our body working. Carbohydrates such as glucose, fructose, starch etc. Are the major source of energy to the human body. They undergo combustion with the help of oxygen that we inhale to form carbon dioxide and water. For example.

$C_5H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O\ (\ell) + \text{energy}$

All combustion reactions are not accompanied by flame. Combustion is basically oxidation accompanied by release of energy.

3.3 (b) Respiration :

Respiration is the most important biochemical reaction which releases energy in the cells. When we breathe in air, oxygen enters our lungs and passes into thousands of smalls air sacs (alveoli). These air sacs occupy a large area of membranes and oxygen diffuses from the membranes into blood. It binds itself to hemoglobin present in red blood cells and is carried to millions of cells in the body. Respiration occurs in these cells and is accompanied by the combustion of glucose producing carbon dioxide and water. Since the reaction is of exothermic nature, the energy released during respiration carry out many cell reactions and also keeps our hart and muscles working. It also provides the desired warmth to the body. Both carbon dioxide and water pas back into the blood and we ultimately breathe them out. Respiration takes place in the cells of all living beings.

Fish takes up oxygen dissolved in water through their gills while plants take up air through small pores (stomata) present in their leaves.

3.3 (c) Harmful Effects of Combustion :

We have discussed the utility of combustion in releasing energy which our body needs to keep warm and working; however, combustion has harmful effects also. The environmental pollution is basically due to combustion. Poisonous gases like carbon monoxide (CO), sulphur dioxide (SO_2) sulphur trioxide (SO_3) and oxide of nitrogen (NO_X) etc. are being released into the atmosphere as a result of variety of combustion reaction which are taking place. They pollute the atmosphere and make our lives miserable. In addition to these, other harmful effects of combustion are corrosion and rancidity. These are briefly discussed.

(i) Corrosion : Corrosion may be defined as the process of slow eating up of the surfaces of certain metals when kept in open for a long time.
Quite often, when we open the bonnet of a car after a long time, we find a deposit around the terminals of the battery. This is an example of corrosion. Black coating on the surface of silver and green layer on the surface of copper are the examples of corrosion. In case of iron, corrosion is called rusting. Rust is a chemical substance brown in colour and is formed by the chemical action of moist air (containing O_2 and H_2O) on iron. It is basically an oxidation reaction and the formula of rust is Fe_2O_3, xH_2O. It is very slow in nature and once started keeps on.
Both corrosion and rusting are very harmful and case damage to the building, Railway tracks, cars and other objects/ materials where metals are used. We quite often hear that an old building has collapsed on its own causing loss of both lives and property. This is on account of the rusting of iron which is used in making the structure particularly the roof.
(ii) Rancidity : Oxidation has damaging effects on food and eatables. When the fats and oils present in butter and margarine are oxidised, they become rancid. As a result, their smell and taste change. They become quite unpleasant. This is known a rancidity. It can be checked in a number of away.

(A) Manufacturer sometimes add certain food additives to the food materials. These are known as antioxidant and check their oxidation.

(B) Keeping food in air tight containers prevents its oxidation.

(C) Refrigeration of food also slows down rancidity because the temperature inside refrigerator is very low and direct contact with air or oxygen is avoided.

(D) Chips manufacturers generally flush their bags with nitrogen before packing so that they may not be oxidised.

DAILY PRACTICE PROBLESM # 3

OBJECTIVE DPP-3.2

1. In the reaction $Mg + Cl_2 \rightarrow MgCl_2$

   Chlorine may be regarded as -

   (A) an oxidising agent (B) a reducing agent

   (C) a catalyst (D) providing an inert medium

2. When the gases sulphur dioxide and hydrogen sulphide react, the reaction is

   $SO_2 + 2H_2S \rightarrow 2H_2O + 3S$

   Here hydrogen sulphide is acting as -

   (A) an oxidising agent (B) a reducing agent

   (C) a dehydrating agent (D) a catalyst

3. Which of the following statements is/are false for oxidation reaction?

   (A) Gain or addition of electronegative radical

   (B) Removal of hydrogen atom.

   (C) Removal or loss of electropositive radical or element

   (D) None of these

4. $CiO + H_2 \rightarrow H_2O + Cu$, reaction is an example of -

   (A) redox reaction (B) synthesis reaction

   (B) neutralisation (D) analysis reaction

5. Which of the following is an example of oxidation reaction ?

(A) $Sn^{+2} - 2e^{-} \rightarrow Sn^{+4}$ (B) $Fe^{+3} + e^{-} \rightarrow Fe^{+2}$

(C) $Cl_2 + 2e^{-} \rightarrow 2Cl$ (D) None of these

6. In the process of burring of magnesium in air, magnesium undergoes -

(A) reduction (B) sublimation (C) oxidation (D) all of these

7. A substance which oxidises itself and reduces other is known as-

(A) an oxidising agent (B) a reducing agent (C) Both of these (D) None of these

8. Oxidation is a process which involves -

(A) addition of oxygen (B) removal of hydrogen

(C) loss of electrons (D) All are correct

9. In the reaction $PbO + C \rightarrow Pb + CO$.

(A) PbO is oxidised

(B) C acts as oxidising agent.

(C) C acts as a reducing agent.

(D) This reaction does not represent a redox reaction.

10. A redox reaction is one in which -

(A) both the substances are reduced.

(B) both the substances are oxidised.

(C) and acid is neutralised by the base.

(D) one substance is oxidised, which the other is reduced.

SUBJECTIVE DPP-3.2

1. Oxidation reaction have some harmful effects. Comment on the sentence.

2. Can oxidation occur without reduction ? Explain

3. Explain the terms oxidation and reduction with examples.

4. What is rancidity? Example with example.

5. What do you mean by corrosion ?

6. Identify the substances that are oxidized and the substances that are reduced in the following reactions -

(a) $ZnO + C \longrightarrow Zn + CO$

(b) $MnO_2 + 4HCl \longrightarrow MnCl_2 + 2H_2O + Cl_2$

(c) $2FeCl_3 + H_2S \longrightarrow 2FeCl_2 + S + 2HCl$

(d) $3Mg + N_2 \longrightarrow Mg_3N_2$

ANSWERS

OBJECTIVE DPP 1.1

Quse.	1	2	3	4	5	6	7	8	9	10
Ans.	C	B	C	A	D	D	C	A	B	C

SUBJECTIVE DPP 1.1

1. (i) $2HgO \longrightarrow 2Hg + O_2$

   (ii) $C_4H_{10} + \frac{13}{2}O_2 \longrightarrow 4CO_2 + 5H_2O$

4. (i) H_2SO_4 (aq) + 2NaOH (aq $\longrightarrow$ Na_2SO_4 (aq) + $2H_2O(\ell)$

   (ii) P_4 (d) + 10 Cl_2 (g) $\longrightarrow$ 4 PCl_5 (g)

5. (i) $ZnCO_3$ (s) $\longrightarrow$ ZnO (s) + Co_2 (g)

   (ii) 2KBr (aq) + BaI_2 (aq) $\longrightarrow$ 2KI (q) + $BaBr_2$ (aq)

OBJECTIVE DPP 2.1

Quse.	1	2	3	4	5	6	7	8	9	10
Ans.	A	B	B	D	C	D	D	B	B	D

OBJECTIVE DPP 2.1

1. (i) Endothermic Reaction
   (ii) Analysis reactions
   (iii) Double displacement reaction
   (iv) Single displacement reaction
   (v) Combination reaction

4. Decomposition reaction

   2AgCl (s) $\longrightarrow$ 2Ag$\downarrow$ + Cl_2 (g)
   (X) grey

6. Cu (s) + $2AgNO_3$ (aq) $\longrightarrow$ Cu $(NO_3)_2$ (aq) + 2Ag (s)

OBJECTIVE DPP 3.1

Quse.	1	2	3	4	5	6	7	8	9	10
Ans.	A	B	D	A	A	C	B	D	C	D

SUBJECTIVE DPP 3.1

6. (a) ZnO is reduced and C is oxidised.
   (b) MnO_2 is reduced and HCl is oxidised.
   (c) $FeCl_3$ is reduced and H_2S is oxidised.
   (d) Mg is oxidised and N_2 is reduced.

Teko CLASSES ....the support

➤ACIDS,BASES AND SALTS➤

4.1 ACIDS :

Substances with sour taste are regarded as avoids. Lemon juice, vinegar, grape fruit juice and spoilt milk etc. taste sour since they are acidic. Many substances can be identified as acids based on their taste but some fo the acids like sulphuric acid have very strong action on the skin which means that they are corrosive in nature. In such case it would be according to modern definition-

An acid may be defined as a substance which release one or more H^+ ions in aqueous solution.

Acids are mostly obtained from natural sources. One the basis of their source avids are of two types -

(a) Mineral acids **(b) Organic acids**

4.1 (a) Mineral Acids :

Acids which are obtained from rocks and minerals are called mineral acids.

4.1 (b) Organic Acids :

Acids which are present in animals and plants are known as organic acids. A list of commonly used acids along with their chemical formula and typical uses, is given below -

Name	Type	Chemical Formula	Where found or used
Carbonic acid	Mineral acid	H_2CO_3	In soft drinks and lends fizz, In stomach as gastric juice, used in tanning industry
Nitric acid	Mineral Acid	HNO_3	Used in the manufacture of explosives. (TNT, Nitroglycerine) and fertilizers (Ammonium nitrate, Calcium nitrate, Purification of Au, Ag.)
Hydrochloric acid	Mineral Acid	HCI	In purification of common salt, in textile industry as bealching agent, to make aqua regia mixture of Hu2HNO3 in ration of 3 : 1
Sulhuric acid	Mineral Acid	H_2SO_4	Commonly used in car batteries, in the manufacture of fertilizers (Ammonium sulphate, super phosphate) detergents etc, in paints, plastics, drugs, in manufacture of artificial silk, in petroleum refining.
Phosphoric acid	Mineral Acid	H3PO4	Used I antirust paints and in fertilizers.
Formic acid	Organic Acid	$HCOOH(CH_2O_2)$	Found in the stings of ants and bees, used in tanning leather, in medicines for treating gout disease of jointly.
Acetic acid	Organic Acid	$CH_3COOH(C_2H_4O_2)$	Fount in vinegar used a solvent in the manufacture of dyes and perfumes.
Lactic acid	Organic Acid	$CH_3CH(OH)COOH(C_3H_6O_3)$	Responsible for souring of milk in curd.
Benzoic acid	Organic Acid	C_6H_5COOH	Used as a food preservation.
Critic acid	Organic Acid	C_6H_8O	Present in lemons, oranges and citrus fruits.

4.1 (c) Chemical Properties of Acids:

1. Action with metals: Dilute acids like dilute HCI and dilute H_2SO_4 react with certain active metals to evolve hydrogen gas.

$2Na(s) + 2HCl\,(dilute) \longrightarrow 2NaCl(aq) + H_2\,(g)$

$Mg(s) + H_2SO_4\,(dilute) \longrightarrow MgSO_4\,(aq) + H_2(g)$

Metals which can displace hydrogen from dilute acids are known as avtive metals. e.g. Na, K, Zn, Fe, Ca, Mg etc.
$Zn(s) + H_2SO_4\,(dilute) \rightarrow ZnSO_4(aq) + H_2(g)$
The active metals which lie above hydrogen in the activity series are electropositive in nature. Their atoms lose electrons to form positive ions and these electrons are accepted by H^+ ions of the acid. As a result, H_2 is evolved.

For e.g.

$Zn(s) \longrightarrow Zn^{2+}\,(aq) + 2e^-$

$ZH^+(aq) + SO_4^{2-}(aq) + 2e^- \longrightarrow H_2\,(g) + SO_4^{2-}\,(aq)$

$Zn(s) + 2H^+\,(aq) \longrightarrow Zn^{++}(aq) + H_2(g)$

2. Action with metal oxides : Acids react with metal oxides to form salt and water. These reactions are mostly carried out upon heating.
For e.g.

$ZnO(s) + 2HCl(aq) \longrightarrow ZNCl_2(aq) + H_2O\,(\ell)$

$MgO(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2O\,(\ell)$

$CuO(s) + 2HCl(dil.) \longrightarrow CuCl_2(aq) + H_2O\,(\ell)$

(Block) (Bluish green)

3. Action with metal carbonates and metal bicarbonates : Both metal carbonates and bicarbonates react with acids to evolve CO_2 gas and form salts.
For e.g.

$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O\,(\ell) + CO_2(g)$

Calcium carbonate Calcium chloride

$2NaHCO_3(s) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2O(aq) + CO_2(g)$

Sodium bicarbonate Sodium sulpahte

4. Action with bases : Acids react with bases to give salts and water.
$HCl + NaOH \longrightarrow NaCl + H_2O$

4.1 (d) Strong and Weak Acids :

(i) Strong acids : Acids which are completely ionised in water are known as strong acids.
For e.g.
Hydrochloric acid (HCl), sulphuric acid (H_2SO_4), nitric acid (HNO_3) etc. are all strong acids.

HCl + Water $\longrightarrow$ $H^+(aq)$ + $Cl^-(aq)$

H_2SO_4 + Water $\longrightarrow$ $2H^+(aq)$ + $SO_4^{2-}(aq)$

(ii) Weak acids: Acids which are weakly ionised in water are known as weak acids.
For e.g.
Carbonic acids (H_2CO_3), phosphoric acid (H_3PO_4), formic acid (HCOOH), acetic acid (CH_3COOH) are weak acids.

CH_3COOH + Water $\longrightarrow$ $CH_3COO^-(aq)$ + H^+ (aq)

In general MINERAL acids are STRONG acids while ORGANIC acids are WEAK acids.

4.2 Base :

Substances with bitter taste and soapy touch are regarded as bases. Since many bases like sodium hydroxide and potassium hydroxide have corrosive action on the skin and can even harm the body, so according to the modern definition -
a base may be defined as a substance capable of releasing one or more OH^- ions in aqueous solution.

4.2 (a) Alkalies :

Some bases like sodium hydroxide and potassium hydroxide are water soluble. These are known as alkalies. Therefore water soluble bases are known as alkalies eg. KOH, NaOH. A list of a few typical bases along with their chemical formulae and uses is given below-

Name	Commercial Name	Chemical Formula	Uses
Sodium hydroxide	Caustic Soda	NaOH	In manufacture of soap, paper, pulp, rayon, refining of petroleum etc.
Potassium hydroxide	Caustic Sba	KHO	In alkaline storage batteries, manufacture of soap, absorbing CO_2 gas etc.
Calcium hydroxide	Slaked lime	$Ca(OH)_2$	In manufacture of bleaching powder softening of hard water etc.
Magnesium hydroxide	Mil of Magnesia	$Mg(OH)_2$	As an antacid to remove acidity from stomach
Aluminum hydroxide	-	$Al(OH)_3$	As foaming agent in fire extinguishers.
Ammonium hydroxide	-	NH_4OH	In removing greases stains from cloths and in cleaning window panes.

4.2 (b) Chemical Properties :

1. Action with metals : Metals like zinc, tin and aluminum react with strong alkalies like NaOH (caustic soda), KOH (caustic potash) to evolve hydrogen gas.

$Zn(s) + 2NaOH(aq) \longrightarrow \underset{\text{Sodium zincate}}{Na_2ZnO_2(aq)} + H_2(g)$

$Sn(s) + 2NaOH(aq) \longrightarrow \underset{\text{Sodium stannite}}{Na_2SnO_2(aq)} + H_2(g)$

$2Al(s) + 2NaOH + 2H_2O \longrightarrow \underset{\text{Sodium meta aluminate}}{2NaAlO_2(aq)} + 3H_2(g)$

2. Action with non-metallic oxides: Acids react with metal oxides, but bases react with oxides of non-metals to form salt and water.

For e.g.

$2NaOH(aq) + CO_2(g) \longrightarrow Na_3CO_3(aq) + H_2O(\ell)$

$Ca(OH)_2(s) + SO_2(g) \longrightarrow CaSO_3(aq) + H_2O(\ell)$

$Ca(OH)_2(s) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(\ell)$

4.2 (c) Strong and Weak Bases :

(i) Strong base : A base contains one or more hydroxyl (OH) groups which it releases in aqueous solution upon ionisation. Bases which are almost completely ionised in water, are known as strong bases.

For e.g.
Sodium hydroxide (NaOH), potassium hydroxide (OH) groups which it releases in aqueous solution upon ionisation. Bases which are almost completely ionised in water, are known as strong bases.

$NaOH(s) + Water \longrightarrow Na^+(aq) + OH^-(aq)$

$KOH(s) + Water \longrightarrow K^+(aq) + OH^-(aq)$

Both NaOH and KOH are deliquescent in nature which means that they absorb moisture from air and get liquefied.

(ii) Weak bases : Bases that are feebly ionised on dissolving in water and reduce a low concentration of hydroxyl ions are called weak bases.

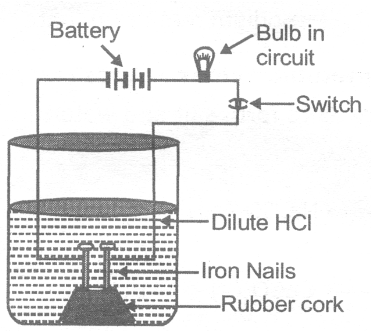
eg. Ca(OH)2, NH_4OH

4.3 CONDUCTING NATURE OF ACID AND BASE SOLUTIONS :

Acids are the substances which contain one or more hydrogen atoms in their molecules which they can release in water as H^+ ions. Similarly, bases are the substances which contain one or more hydroxyl groups in their molecules which they an release in water as OH^- ions. Since the ions are the carries of charge therefore, the aqueous solutions of both acids and bases are conductors of electricity.

Experiment :

In a glass beaker, take a dilute solution of hydrochloric acid (HCl). Fix two small nails of iron in a rubber cork in the beaker as shown in the figure. Connect the nails to the terminals of a 6 volt battery through a bulb. Switch on the current and bulb will start glowing. This shows that the electric current has passed through the acid solution. As the current is carried by the movement of ions, this shows that is solution HCl has ionised to give H^+ and Cl^- ions. Current will also be in a position to pass if the beaker contains in it dilute H_2SO_4 (H^+ ions are released in aqueous solution). Similarly, aqueous solutions containing NaOH or KOH will also be conducting due to release of OH^- ions.



Bulb will not glow if glucose ($C_6H_{12}O_6$) or ethyl alcohol (C_2H_6O) solution is kept in the beaker. This means that both of them will not give any ions in solution.

4.4 COMPARISON BETWEEN PROPERTIES OF ACIDS AND BASES :

Acids	Bases
1. Sour in taste. 2. Change colours of indicators et. Litmus turns from blue to red, phenolphthalein remains colourless. 3. Shows electrolytic conductivity in aqueous solution. 4. Acidic properties disappear when reacts with bases (Neutralisation). 5. Acids decompose carbonate salts.	1. Bitterness in taste. 2. Change colours of indicators eg, litmus turns from red to blue, phenolphthalein furns from colourless to pink. 3. Shows electrolytic conductivity in aqueous solutions. 4. Basic properties disappear when reacts with acids (Neutralisation). 5. No decomposition of carbonate salts by bases.

4.5 ROLE OF WATER IN THE INISATION OF ACIDS AND BASES :

Substances can act as acids and bases only in the presence of water in aqueous solution. In dry state which is also called anhydrous state, these characters cannot be shown Actually, water helps in the ionisation of acids or base by separating the ions. This is also known as dissociation and is explained on the basis of a theory called Arrhenius theory of acids and bases.

In the dry state, hydrochloric acid is known as hydrogen chloride gas i.e. HCl(g). It is not in the position to give any H^+ ions. Therefore, the acidic character is not shown. Now, let us pass the gas through water taken in a beaker with the help of glass pipe. H_2O molecules are of polar nature which means that they have partial negative charge (δ^+) on oxygen atom and partial positive charge (δ^-) on hydrogen atoms. They will try to form a sort of envelope around the hydrogen atoms as well as chlorine atoms present in the acid and thus help in their separation as ions. These ions are said to be hydrated ions.

$$HCl(g) + Water \longrightarrow \underset{\text{(Hydrated ions)}}{H^+ (aq) + Cl^- (aq)}$$

The electrical current is carried through these ions. The same applied to other acids as well as bases. Thus we conclude that -

(i) acids can release H^+ ions only in aqueous solution.

(ii) base can release OH^- ions only in aqueous solution.

(iii) hydration helps in the release of ions from acids and bases.

4.6 DILUTION OF ACIDS AND BASES :

Acids and bases are mostly water soluble and can be diluted by adding the required amount of water. With the addition of water the amount of acid or base per unit volume decrease and dilution occurs. The process is generally exothermic in nature. A concentrated acid like sulphuric acid or nitric acid is to be diluted with water. Acid should be added dropwise to water taken in the container with constant stirring.

DAILY PRACTICE PROBLEMS # 4

OBJECTIVE DPP - 4.1

1. The acid used in making of vinegar is -
(A) Formic acid (B) Acetic acid (C) Sulphuric acid (D) Nitric acid

2. Common name of H_2SO_4 is-
(A) Oil of vitriol (B) Muriatic acid (C) Blue vitriol (D) Green vitriol

3. $CuO + (X) \longrightarrow CuSO_4 + H_2O$. Here (X) is-
(A) $CuSO_4$ (B) HCl (C) H_2SO_4 (D) HNO_3

4. Which of the following is the weakest base ?
(A) NaOH (B) NH_4OH (C) KOH (D) $Ca(OH)_2$

5. Reaction of an acid with a base is known as-
(A) decomposition (B) combination (C) redox reaction (D) neutralization

6. When CO_2 is passed through lime water, it turns milky; The milkiness is due to the formation of -
(A) $CaCO_3$ (B $Ca(OH)_2$ (C) H_2O (D) CO_2

7. Caustic soda is the common name for-
(A) $Mg(OH)_2$ (B) KOH (C) $Ca(OH)_2$ (D) NaOH

8. Antacids contain -
(A) Weak base (B) Weak acid (C) Strong base (D) Strong acid

9. Calcium hydroxide (slaked lime) is used in -
(A) Plastics and dyes (B) Fertilizers (C) Antacids (D) White washing

10. Acids gives -
(A) H^+ in water (B) OH^- in water (C) Both (A) & (B) (D) None of these

11. H_2CO_3 is a -
(A) strong acid (B) weak acid (C) strong base (D) weak base

SUBJECTIVE DPP- 4.2

1. Equal amounts of calcium are taken in test tubes (A) and (B). Hydrochloric acid (CHI) is added to test tube (A) while acetic acid (CH_3COOH) is added to test tube (B). In which case, fizzing occurs more vigorously and why ?

2. Give the name of two mineral acids and their uses.

3. What effect does concentration of H^+ (aq) have on acidic nature of the solution?

4. What do you understand by organic acids? Give the name of the organic acids and their sources.

5. Which gas is usually liberated when an acid reacts with metal? Illustrate with an example how will you test the presence of the gas?

➤ACIDS, BASES AND SALTS⮜

5.1 INDICATORS:

Indicator indicated the nature of particular solution whether acidic, basic or neutral. Apart from this, indicator also represents the change in nature of the solution from acidic to basic and vice versa. Indicators are basically coloured organic substances extracted from different plants. A few common acid base indicators are

5.1 (a) Litmus :

Litmus is a purple dye which is extracted from 'lichen' a plant belonging to variety Thallophytic. It can also be applied on paper in the form of strips and is available as blue and red strips. A blue litmus strip, when dipped in an acid solution acquires red colour. Similarly a red strip when dipped in a base solution becomes blue.

5.1 (b) Phenolphthalein :

It is also an organic dye and acidic in nature. In neutral or acidic solution, it remains colourless while in the basic solution, the colour of indicator changes to pink.

5.1 (c) Methyl Orange :

Methyl orange is an orange coloured dye (yellow) and basis in nature. In the acidic medium the colour of indicator becomes red and in the basic or natural medium, it colour remains unchanged.

5.1 (d) Red Cabbage Juice :

It is purple in colour in natural medium and turns red or pink in the acidic medium. In the basic or alkaline medium, its colour changes to green.

5.1 (e) Turmeric Juice :

It is yellow in colour and remains as such in the neutral and acidic medium. In the basic medium its colour becomes reddish or deep brown.

Sample	Blue litmus solution	Red litmus solution	Phenolphthalein	Methyl orange
HCl	Changes to red	No colour change	Remains colourless	Changes to red
HNO_3	Changes to red	No colour change	Remains colourless	Changes to red
NaOH	No colour change	Changes to blue	Changes to light pink	No changes in colour
KOH	No colour change	Changes to blue	Changes to light pink	No changes in colour

Litmus is obtained from LICHEN plant.

5.2 NEUTRALISATION :

It may be defined as a reaction between acid and base present in aqueous solution to form salt and water.

$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(\ell)$

Basically neutralisation is the combination between H^+ ions of the acid with OH^- ions of the base to form H_2O.

For e.g.

$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + H_2O(\ell)$

$H^+(aq) + OH^-(aq) \longrightarrow H_2O(\ell)$

Neutralisation reaction involving an acid and base is of exothermic nature. Heat is evolved in all naturalisation reactions. If both acid and base are strong, the value of heat energy evolved remains same irrespective of their nature.

For e.g.

$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O(\ell) + 57.1\ KJ$

(Strong acid)	(Strong base)

$HNO_3\ (aq) + KOH(aq) \longrightarrow KNO_3(aq) + H_3O(\ell) + 57.1\ J$

(Strong acid)	(Strong base)

Strong acids and strong bases are completely ionised of their own in the solution. No energy is needed for their ionisation. Since the action of base and anion of acid on both sides of the equation cancels out completely, the heat evolved is given by the following reaction -

$H^+\ (aq) + OH^-\ (aq) \longrightarrow H_2O(\ell) + 57.1\ KJ$

5.3 APPLICATIONS OF NEUTRALISATION :

(i) People particularly of old age suffer from acidity problems in the stomach which is caused mainly due to release of excessive gastric juices containing HCl. The acidity is neutralised by antacid tablets which contain sodium hydrogen carbonate (baking soda), magnesium hydroxide etc.

(ii) The sting of bees and ants contain formic acid. Its corrosive and poisonous effect can be neutralised by rubbing soap which contains NaOH (an alkali).

(iii) The stings of wasps contain an alkali and its poisonous effect can be neutralised by an acid like acetic acid (present in vinegar).

(iv) Farmers generally neutralise the effect of acidity in the soil caused by acid rain by adding slaked lime (Calcium hydroxide) to the soil.

5.4 pH SCALE :

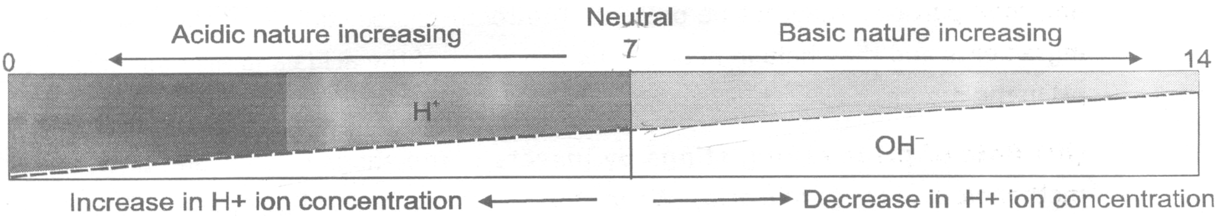
A scale for measuring hydrogen ion concentration in a solution called pH scale, has been developed by S.P.L. sorrensen. The P in pH stands for potenz' in German meaning power. On the pH scale we can measure pH from O (very acidic) to 14 (very alkaline). pH should be thought of simply as a number which indicates the acidic or basic nature of solution. Higher the hydrogen ion concentration, Lower is the pH scale.

Characteristic of pH scale are -

(i) For acidic solution, pH < 7

(ii) For alkaline solution, pH > 7

(iii) For neutral solution, pH = 7



5.4 (a) Universal Indicator Papers for pH Values :

Indicators like litmus, phenolphthalein and methyl orange are used in predicting the acidic and basic characters of the solutions. However universal indicator papers have been developed to predict the pH of different solutions. Such papers represent specified colours for different concentrations in terms of pH values.

The exact pH of the solution can be measured with the help of pH meter which gives instant reading and it can be relied upon.

pH values of a few common solutions are given below -

Solution	Approximate pH	Solution	Approximate pH
Gastric juices	1.0 – 3.0	Pure water	7.0
Lemon juices	2.2 - 2.4	Blood	7.36 – 7.42
Vinegar	3.0	Baking soda solution	8.4
Bear	4.0 – 5.0	Sea water	9.0
Tomato juice	4.1	Washing soda solution	10.5
Coffee	4.5 – 5.5	Lime water	12.0
Acid rain	5.6	House hold ammonia	11.9
Milk	6.5	Sodium hydroxide	14.0
Saliva	6.5 – 7.5		

5.4 (b) Significance of pH in daily life :

(i) pH i our digestive system : Dilute hydrochloric acid produced in our stomach helps in the digestion of food. However, excess of acid causes indigestion and leads to pain as well as irritation. The pH of the digestive system in the stomach will decrease. The excessive acid can be neutralised with the help of antacid which are recommended by the doctors. Actually, these are group of compounds (basic in nature) and have hardly and side effects. A very popular antacid is 'Milk of Magnesia' which is insoluble magnesium hydroxide. Aluminum hydroxide and sodium hydrogen carbonate can also be used for the same purpose. These antacids will bring the pH of the system back to its normal value. The pH of human blood varies between 7.36 to 7.42. it is maintained by the soluble bicarbonates and carbonic acid present in the blood. These are known as

buffers.

(ii) pH change leads to tooth decay : The white enamel coating on our teeth is of insoluble calcium phosphate which is quite hard. It is not affected by water. However, when the pH in the mouth falls below 5.5 the enamel gets corroded. Water will have a direct access to the roots and decay of teeth will occur. The bacteria present in the mouth break down the sugar that we eat in one form or the other to acids, Lactic acid is one these. The formation of these acids causes decrease in pH. It is therefore advisable to avoid eating surgery foods and also to keep the mouth clean so that sugar and food particles may not be present. The tooth pastes contain in them some basic ingredients and they help in neutralising the effect of the acids and also increasing the pH in the mouth.

(iii) Role of pH in curing stings by insects: The stings of bees and ants contain methanoic acid (or formic acid). When stung, they cause lot of pain and irritation. The cure is in rubbing the affected area with soap. Sodium hydroxide present in the soap neutralises acid injected in the body and thus brings the pH back to its original level bringing relief to the person who has been stung. Similarly, the effect of stings by wasps containing alkali is neutralised by the application of vinegar which is ethanoic acid (or acetic acid)

(iv) Soil pH and plant growth : The growth of plants in a particular soil is also related to its pH. Actually, different plants prefer different pH range for their growth. it is therefore, quite important to provide the soil with proper pH for their healthy growth. Soils with high iron minerals or with vegetation tend to become acidic. This soil pH can reach as lows as 4. The acidic effect can be neutralised by 'liming the soil' which is carried by adding calcium hydroxide. These are all basic in nature and have neutralising effect. Similarly, the soil with excess of lime stone or chalk is usually alkaline. Sometimes, its pH reaches as high as 8.3 and is quite harmful for the plant growth. In order to reduce the alkaline effect, it is better to add some decaying organic matter (compost or manure). The soil pH is also affected by the acid rain and the use of fertilizers. Therefore soil treatment is quite essential.

DAILY PRACTICE PROBLEMS # 5

OBJECTIVE DPP-5.1

1. A solution turns red litmus blue. Its pH is likely to be-
(A) 2 (B) 4 (C) 7 (D) 10

2. If pH of any solution is equal to zero then solution will be-
(A) acidic (B) basic (C) neutral (D) none of these

3. Methyl orange is -
(A) an acidic indicator (B) a basic indicator (C) a neutral indicator (D) none of these

4. pH of Blood is-
(A) 6.4 (B) 7.4 (C) 4.7 (D) 6.4

5. If ph of solution is 13, means that it is-
(A) weakly acidic (B) weakly basic (C) strongly acidic (D) strongly basic

6. Which is a base and not an alkali ?
(A) NaOH (B) KOH (C) $Fe(OH)_3$ (D) None is true

7. Energy released in neutralisation reaction which occurs between strong acid and strong base is-
(A)57.8 kJ (B) 57.1 kJ (C) hNO_3 (D) $H_2C_2O_4$

9. A solution has pH 9. On dilution the pH value
(A) decreases (B) increases (C) remain same (D) none of these

SUBJECTIVE DPP-5.2

1. Five solutions A,B,C,D and E when tested with universal indicator shows pH as 5, 3, 13, 7 and 9 respectively. Which solution is -

   (a) neutral.

   (b) strongly alkaline.

   (c) strongly acidic.

   (d) weakly alkaline.
   (e) weakly acidic.
   Arrange the pH in decreasing order of H^+ ion concentration.

2. What will you observe when-

   (i) red litmus paper is introduced into a solution of sodium sulphate ?

   (ii) methyl orange is added to dilute hydrochloric acid ?

   (iii) a drop of phenolphthalein is added to solution of lime water ?

   (iv) blue litmus is introduced into a solution of ferric chloride ?

3. Give two applications of pH in our daily life.

4. Explain why ?

   (i) Aqueous solution of sodium acetate has pH more than 7.

   (ii) Aqueous solution of copper sulphate has pH less than 7.

   (iii) Aqueous solution of Potassium nitrate has pH value 7.

ACIDS, BASES AND SALTS

6.1 SALTS :

A substance formed by neutralization of an acid with a base is called a salt.

For e.g.

$Ca\ (OH)_2 + H_2SO_4 \rightarrow CaSO_4 + H_2O$

$2Ca(OH)_2 + 4HNO_3 \rightarrow 2Ci(NO_3)_2 + 2H_2O$

$NaOH + HCl \rightarrow NaCl + H_2O$

6.2 CLASIFFICATION OF SALTS :

Salts have been classified on the basis of chemical formulae as well as pH values.

6.2 (a) Classification Based on Chemical Formulae :

(i) Normal salts : A normal salt is the one which does not contain any ionsable hydrogen atom or hydroxyl group. This means that it has been formed by the complete neutralisation of an acid by a base.
For e.g. NaCl, KCl, $NaNO_2$, $K_2\ SO_4$ etc.

(ii) Acidic salts : an acidic salt still contains some replaceable hydrogen atoms, This means that the neutralisation of acid by the base is no complete. For example, sodium hydrogen sulphate ($NaHCO_4$), sodium hydrogen carbonate ($NaHCl_3$) etc.

(iii) Basic salts : A basic salt still contains some replaceable hydroxyl groups. This means that the neutralisation of base by the acid is not complete. For example, basic lead nitrate Pb (OH) NO_3. basic lead chloride, Pb(OH)Cl etc.

6.2 (b) Classification Based on pH Values :

Salts are formed by the reaction between acids and bases. Depending upon the nature of the acids and bases or upon the pH values, the salt solutions are of three types.

(i) Neutral salt solutions : Salt solutions of strong acids and strong bases are neutral and have pH equal to 7. They do not change the colour of litmus solution.
For e.g. : NaCl, $NaNO_3$, Na_2SO_4 etc.

(ii) Acidic salt solutions : Salt solutions of strong acids and weak bases are of acidic nature and have pH less than 7. They change the colour of blue litmus solution to red.

For e.g. $(NH_4)_2SO_4$, NH_4Cl etc.
In both these salts, the base NH_4OH is weak while the acids H_2SO_4 and HCl are strong.

(iii) Basic salt solutions : Salt solutions of strong bases and weak acids are of basic nature and have pH more than 7. They change the colour of red litmus solution to blue.

For e.g. Na_2CO_3,K_3PO_4 etc.
In both the salts, bases NaOH and KOH are strong while the acids H_2CO_3 and H_3PO_4 are weak.

6.3 USES OF SALTS :

(i) As a table salt,

(ii) In the manufacture of butter and cheese.

(iii) In leather Industry.

(iv) In the manufacturing of washing soda and baking soda.

(v) For the preparation of sodium hydroxide by electrolysis of brine.

(vi) Rock salt is spread on ice to melt it in cold countries.

SOME IMPORTANT CHEMICAL COMPOUNDS :

6.4 SODIUM CHLORIDE - COMMON SALT (TABLE SALT) :

Sodium chloride (NaCl) also called common salt or table salt is the most essential part of our diet. Chemically it is formed by the reaction between solutions of sodium hydroxide and hydrochloric acid. Sea water is the major source of sodium chloride where it is present in disserved form along with other soluble salts such as chlorides and sulphates of calcium and magnesium. it is separated by some suitable methods. Deposits of the salts are found in different part of the world and is known as rock salt. When pure, it is a white crystalline solid, However, it is often brown due to the presence of impurities.

6.4 (a) Uses :

(i) Essential for life : Sodium chloride is quite essential for life. Biologically, it has a number of function to perform such as in muscle contraction, in conduction of nerve impulse in the nervous system and is also converted in hydrochloric acid which helps in the digestion of food in the stomach. When we sweat, there is loss of sodium chloride along with water. It leads to muscle cramps. Its loss has to be compensated suitably by giving certain salt preparations to the patient. Electrol powder is an important substitute of common salt.

(ii) Raw material for chemical: Sodium chloride is also a very useful raw material for different chemical. A few out of these are hydrochloric acid (HCl), washing soda ($Na_2CO_3.10H_2O$), baking soda ($NaHCO_3$) etc. Upon electrolysis of a strong solution of the salt (brine), sodium hydroxide, chlorine and hydrogen are obtained. Apart from these, it is used in leather industry for the leather tanning. In severe cold, rock salt is spread on icy roads to melt ice. it is also used as fertilizer for sugar beet.

6.4 (b) Electrolysis of aqueous solution of NaCl :

$$2NaCO\,(s) + 2H_2O\,(\ell) \xrightarrow{\text{Electrolysis}} 2NaOH(aq) + Cl_2(g) + H_2(g)$$

reaction takes place in two steps

(i) $2Cl^- \longrightarrow Cl_2(g) + 2e^-$ (anode reaction)

(ii) $2H_2O + 2e^- \longrightarrow H_2 + OH^-$ (cathode reaction)

6.5 WASHING SODA :

Chemical name :
Sodium carbonate decahydrate
Chemical formula : Na_2CO_3, $10H_2O$

6.5 (a) Recrystallization of sodium carbonate:

Sodium carbonate is recrystallized by dissolving in water to get washing soda it is a basic salt.

$$\underset{\text{Sodium carbonate}}{Na_2CO_3} + 10H_2O \longrightarrow \underset{\text{Washing soda}}{Na_2CO_3, 10H_2O}$$

6.5 (b) Uses :

(i) It is used as cleansing agent for domestic purposes.

(ii) It is used in softening hard water and controlling the pH of water.

(iii) It is used in manufacture of glass.

(iv) Due to its detergent properties, it is used as a constituent of several dry soap powders.

(v) It also finds use in photography, textile and paper industries etc.

(vi) It is used in the manufacture of borax ($Na_2B_4O_7$. $10H_2O$)

6.6 BAKING SODA :

Baking soda is sodium hydrogen carbonate or sodium bicarbonate ($NaCHO_3$).

6.6 (a) Preparation :

It is obtained as an intermediate product in the preparation of sodium carbonate by Solvay process. In this process, a saturated solution of sodium chloride in water is saturated with ammonia and then carbon dioxide gas is passed into the liquid. Sodium chloride is converted into sodium bicarbonate which, being less soluble, separates out from the solution.

$$2NH_3\,(g) + H_2O\,(\ell) + CO_2\,(g) \longrightarrow (NH_4)_2\,CO_3(aq)$$

$$(NH_4)_2CO_3\,(aq) + 2NaCl(aq) \longrightarrow Na_2CO_3\,(aq) + 2NH_4Cl(aq)$$

$$Na_2CO_3(aq) + H_2O\,(\ell) + CO_2\,(g) \longrightarrow 2NaHCO_3(s)$$

6.6 (b) Properties :

(i) It is a white, crystalline substance that forms an alkaline solution with water. The aqueous solution of sodium bicarbonate is neutral to methyl orange but gives pink colour with phenolphthalein orange. (Phenolphthalein and methyl orange are dyes used as acid-base indicators.)

(ii) When heated above 543 K, it is converted into sodium carbonate.

$$2NaHCO_3\,(s) \longrightarrow Na_2CO_3\,(s) + CO_2\,(g) + H_2O\,(\ell)$$

6.6 (c) Uses :

(i) It is used in the manufacture of baking powder. Baking powder is a mixture or potassium hydrogen tartar ate and sodium bicarbonate. During the preparation of bread the evolution of carbon dioxide causes bread the evolution of carbon dioxide causes bread to rise (swell).

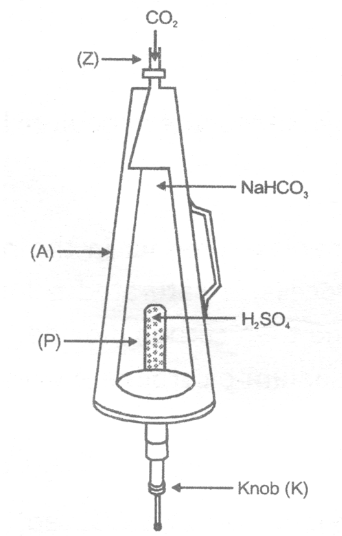
$$\begin{array}{l} CH(OH)COOK \\ | \\ CH(OH)COOH \end{array} + NaHCO_3 \longrightarrow \begin{array}{l} CH(OH)COOK \\ | \\ CH(OH)COONa \end{array} + CO_2 + H_2O$$

(ii) It is largely used in the treatment of acid spillage and in medicine as soda bicarb, which acts as an antacid.

(iii) It is an important chemical in th textile, tanning, paper and ceramic industries.

(iv) It is also used in a particular type of fire extinguishers. The following diagram shows a fire extinguisher that uses $NaHCl_3$ and H_2SO_4 to produce CO_2 gas. The extinguisher consists of a conical metallic container (A) with a nozzle (Z) at one end. A strong solution of $NaHCO_3$ is kept in the container. A glass ampoule (P) containing H_2SO_4 is attached to a knob (K) and placed inside the $NaHCO_3$ solution. The ampoule can be broken by hitting the knob. As soon as the acid comes in contact with the $NaHCO_3$ solution, CO_2 gas is formed. When enough pressure in built up inside the container, CO_2 gas rushes out through the nozzle (A). Since CO_2 does not support combustion, a small fire can be put out by pointing the nozzle towards the fire. The gas is produced according to the following reaction.

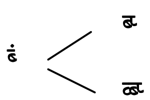
$2NaHCO_3\,(aq) + H_3SO_4\,(aq) \longrightarrow Na_2SO_4\,(aq) + 2H_2O\,(\ell) + 2CO_2(g)$



Fire Extinguisher

6.7 BLEACHING POWDER :

Bleaching powder is commercially called 'chloride of lime or' chlorinated lime'. It is principally calcium oxychloride having the following formula :



Bleaching powder is prepared by passing chlorine over slaked lime at 313 K.

$Ca(OH)32\ (aq) + Cl_2(g) \xrightarrow{313\ K} Ca(OCl)Cl\ (s) + H_2O(g)$

Slaked lime Bleaching powder

Actually beaching powder is not a compound but a mixture of compounds : $CaOCl_2$, $4H_2O$, $CaCl_2$. $Ca(OH)_2$. H_2O

6.7 (a) Uses :

(i) It is commonly used as a bleaching agent in paper and textile industries.

(ii) It is also used for disinfecting water to make water free from germs.

(iii) It is used to prepare chloroform.

(iv) It is also used to make wool shrink-proof.

6.8 PLASTER OF PARIS :

6.8 (a) Preparation :

It is prepared by heating gypsum ($CaSO_4 . 2H_2O$) at about
373 k in large seel pots with mechanical stirrer ,
or in a revolving furnace .

$$\underset{\text{Gypsum}}{2(CaSO_4,2H_2O)} \xrightarrow{373\text{ K}} \underset{\text{Plaster of Parries}}{(CaSO_4)_2, H_2O + 3H_2O}$$

$$\text{or } CaSO_4, 2H_2O \longrightarrow CaSO_4, \frac{1}{2}H_2O + \frac{3}{2}H_2O$$

The temperature is carefully controlled, as at higher temperature
gypsum is fully dehydrated . The properties of dehydrated gypsum are completely
different from those of plaster of Paris.

6.8 (b) Properties :

(i) Action with water : When it is dissolved in water , it gets crystallized and forms gypsum

$$CaSO_4, \frac{1}{2}H_2O + \frac{3}{2}H_2O \longrightarrow CaSO_4,2H_2O$$

6.8 (c) Uses :

When finely powered Plaster of Parries is mixed with water and made into a paste, it quickly sets into a hard mass. In the process, its volume also increases slightly. These properties find a number of uses. Addition of water turns Plaster of Parries back into gypsum.

(i) It is used in the laboratories for sealing gaps where airtight arrangement is required.

(ii) It is also used for making toys, cosmetic and casts of statues.

(iii) It is used as a cast for setting broken bones.

(iv) It also find use in making moulds in pottery.

(iv) It is also used for making surfaces smooth and for making designs on walls and ceilings.

6.9 HYDRATED SATLS - SALTS CONTAINING WATER OF CRYSTALLISATION:

Certain salts contain definite amount of some H_2O molecules loosely attached to their own molecules. These are known as hydrated salts and are of crystalline nature. The molecules of H_2O present are known as 'water of crystallisation'.

In colourd crystalline and hydrated salts, the molecules of water of crystallisation also account for their characteristic colours. Thus, upon heating of hydrated salt, its colour changes since molecules of water of crystallisation are removed and the salt becomes anhydrous, For example, take a few crystals of blue vitriol i.e. hydrated copper sulphate in a dry test tube or boiling tube. Heat the tube from below. The salt will change to a white anhydrous powder and water droplet will appear on the walls of the tube. Cool the tube and add a few droops of water again. The white anhydrous powder will again acquire blue colour.

$$CuSO_4.\,5H_2O \xrightarrow{\Delta} CuSO_4 + 5H_2O$$

Copper sulphate	Copper sulphate
(Hydrated)	(Anhydrous)

DAILY PRACTICE PROBLEMS # 6

OBJECTIVE DPP-6.1

1. A salt derived from strong acid and weak base will dissolve in water to give a solution which is -
(A) acidic (B) basic (C) neutral (D) none of these

2. Materials used in the manufacture of bleaching powder are -
(A) lime stone and chlorine (B) quick lime and chlorine
(C) slaked lime and HCl (D) slaked lime and chlorine

3. Bleaching powder gives smell of chlorine because it-
(A) is unstable (C) gives chlorine on exposure to atmosphere
(C) is mixture of chlorine and slaked lime (D) contains excess of chlorine

4. Baking powder contains, baking soda and-
(A) potassium hydrogen tartarate (B) calcium bicarbonate
(C) sodium carbonate (D) vinegar

5. Plaster of pairs is made from-
(A) lime stone (B) slaked lime (C) quick lime (D) gypsum

6. Setting of plaster of Paris takes place due to-
(A) oxidation (B) reduction (C) dehydration (D) hydration

7. Chemical formula of baking soda is-
(A) $mGSO_4$ (B) Na_2CO_3 (C) $NaHCO_3$ (D) $MgCO_3$

8. The chemical name of marble is -
(A) calcium carbonate (B) Magnesium carbonate
(C) calcium chloride (D) calcium sulphate

9. Washing soda has the formula -
(A) $Na_2\ CO_3,\ 7H_2O$ (B) $Na_2CO_3,\ 10H_2O$ (C) Na_2CO_3,H_2O (D) Na_2CO_3

10. The raw materials required for the manufacture of $NaHCO_3$ by Solvay process are -
(A) $CaCl_2,\ (NH_4)_2\ CO_3,\ NH_3$ (B) $NH_4Ci,NaCl_2Ca(OH)_2$
(C) $NaCl_2,(NH_4)_2CO_3,NH_3$ (D) $NaCO,NH_3CaCO_3,H_2O$

11. Plaster of Paries hardens by-
(A) giving off CO_2. (B) changing into $CaCO_3$.
(C) combining with water (D) giving out water.

12. The difference in number of water molecules in gypsum and plaster of paris is-
(A) 5/2 (B) 2 (C) ½ (D) 3/2

SUBJECTIVE DPP-6.2

1. Give chemical names of the following compounds. Also state one use in each case.
(i) Washing soda (ii) Baking soda (iii) Bleaching powder

2. Explain why-
(i0 common salt becomes sticky during the rainy season ?

(ii) blue vitriol changes to white upon heating ?

(iii) anhydrous calcium chloride is used in desiccators ?

(iv) if a bottle full of concentrated sulphuric acid is left open in the atmosphere by accident the acid starts flowing out of the bottle of its own ?

3. How will you prepare the following ? Give chemical reactions also.

(i) Plaster of Paris from Gypsum.

(ii) Bleaching powder from slaked lime.

(iii) Baking soda from brine.

ANSWERS

OBJECTIVE DPP - 4.1

Ques.	1	2	3	4	5	6	7	8	9	10	11
Ans	B	A	C	B	D	A	D	A	D	A	B

OBJECTIVE DPP- 5.1

Ques.	1	2	3	4	5	6	7	8	9
Ans	D	A	B	B	D	D	B	C	A

OBJECTIVE DPP -6.1

Ques.	1	2	3	4	5	6	7	8	9	10	11	12
Ans	A	D	B	A	D	D	C	A	B	D	C	D

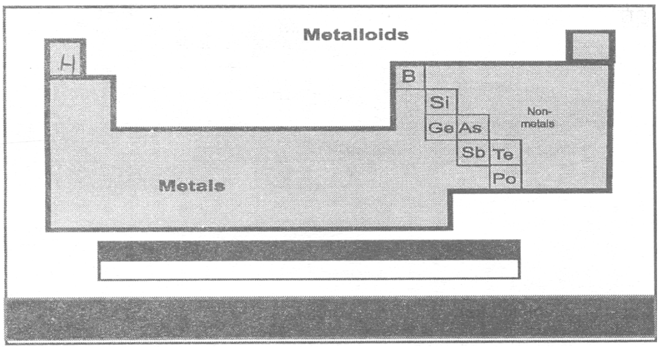
METALS AND NON-METALS

7.1 INTRODUCTION:

There are 118 chemical elements known at present. One the basis of their properties, all these elements can be broadly divided into two main groups: Metals and Non-Metals. A majority of the known elements are metals. All the metals are solids, except mercury, which is a liquid metal. There are 22 non-metals, out of which, 10 non-metals are solids, one non-metal (bromine) is liquid and the remaining 11 non-metals are gases.

7.2 POSITION OF METALS AND NON-METALS IN THE PERIODIC TABLE :

The metals are placed on the left hand side and in the centre of the periodic table. One the other hand, the non-metals are placed on the right hand side of the periodic table. This has been shown in the figure. It may be noted that hydrogen (H) is an exception because it is non-metal but is placed on the left hand side of the periodic table.
Metals and non-metals are separated from each other in the periodic table by a zig-zag line. The elements close to zig-zag line show properties of both the metals and the non-metals. They show some properties of metals and some properties of non-metals. These are called metalloids. The common examples of metalloids are boron (B), silicon (Si), germanium (Ge), arsenic (As), antimony (Sb), tellurium (Te) and polonium (Pi).



In general, the metallic character decreases on going from left to right side in the periodic table. However, on going down the group, the metallic character increases.


Store in your memory

The elements at the extreme left of the periodic table are most metallic and those on the right are least metallic or non-metallic.

7.3 GENERAL PROPERTIES OF METALS AND NON-METALS :

7.3 (a) Electronic Configuration of Metals :

The atoms of metals have 1 to 3 electrons in their outermost shells. For example, all the alkali metals have one electron in their outermost shells (lithium 2, 1; sodium 2,8,1: potassium 2,8,8,1 etc.)
Sodium, magnesium and aluminum are metals having 1,2 and 3 electrons respectively in their valence shells. Similarly, other metals have 1 to 3 electron in their outermost shells.

It may be noted that hydrogen and helium are exception because hydrogen is a non-metal having only electron in the outermost shell (K shell) of its atom and helium is also a non-metal having 2 electron in the outermost shell (K shell).

7.3 (b) Physical Properties of Metals:

The important physical properties of metals are discussed below:

(i) Metals are solids at room temperature: All metals (except mercury) are solids at room temperature.

MERCURY is a liquid at room temperature.

(ii) Metals are malleable: metals are generally malleable. Malleability means that the metals can be beaten with a hammer into very thin sheets without breaking. Gold and silver are among the best malleable metals. Aluminum and copper re also highly malleable metals.

(iii) Metals are ductile : It means that metals can be drawn (stretched) into this wires. Gold and silver are the most ductile metals. Copper and aluminum are also very ductile, and therefore, these can be drawn into this wires which are used in electrical wiring.

(iv) Metals are good conductors of heat and electricity: All metals are good conductors of heat. The conduction of heat is called thermal conductivity. Silver is the best conductor of heat. Copper and aluminum are also good conductors of heat and therefore, they are used for making household utensil. Lead is the poorest conductor of heat. Mercury metal is also poor conductor of heat.
Metals are also good conductors of electricity. The electrical and thermal conductivities of metals are due to the presence of free electrons in them. Among all the metals, silver is the best conductor electricity. Copper and aluminum are the next best conductors of electricity. Since silver is expensive, therefore, copper and aluminum are commonly used for making electric wires.

(v) Metals are lustrous and can be polished: Most of the metals have shine and they can be polished. The shining appearance of metals is also known as metallic lustre. For example, gold, silver and copper metals have metallic lustre.

Store in your memory

SILVER is best conductor of heat and electricity.

(vi) Metals have high densities : Most of the metals are heavy and have high densities. For example, the density of mercury metal if very high (13.6 g cm^{-3}). However, there are some exceptions. Sodium, potassium, magnesium and aluminum have low densities. Densities of metals are generally proportional to their atomic masses. The smaller the metal atom, the smaller it its density.

(vii) Metals are hard : Most of the metals are hard. But all metals are not equally hard. Metals like iron, copper, aluminum etc. are quite hard. They cannot be cut with a knife. Sodium and potassium are common exceptions which are soft and can be easily cut with a knife.

(viii) Metals have high melting and boiling points : Most of the metals (except sodium and potassium) have high melting and boiling points.

Tungsten has highest melting point (3410^0C) among all the metals.

(ix) Metals are rigid : Most of the metals are **rigid** and they have **high** tensile strength.

(x) Metals are sonorous: Most of the metals are **sonorous** i.e., they make sound when hit with an object.

7.3 (c) Electronic Configuration of Non-Metals :

The atoms and non-metals have usually 4 to 8 electrons in their outermost shells. For example, Carbon (At. No6), Nitrogen (At. No. 7), Oxygen (At. No.8), Fluorine (At. No. 9) and Neon (At. No. 10) have respectively 4,5,6,7,8 electrons in their outermost shells.

7.3 (d) Physical Properties Of Non-Metals:

The important physical properties of non-metals are listed below :

(i) Non-metals are brittle.

(ii) Non- metals are not ductile.

(iii) Non-metals are bad conductor of heat and electricity. (Exception: Graphite is a good conductor because of the presence of free electrons.)

(iv) Non-metals are not lustrous and cannot be polished. (Exception: Graphite and Iodine are lustrous non-metals.)

(v) Non-metals may be solid, liquid, or gases at room temperature.

(vi) Non-metals are generally soft. (Exception: Diamond, an allotropic from of non-metal Carbon, is the hardest natural substance known).

(vii) Non-metals have generally low melting and boiling points. (Exception: Graphite another allotropic form of Carbon, has a melting point of about 3730^0C).

(viii) Non-metals have low densities. (Exception : Iodine has high density).


Store in your memory

Graphite is a good conductor of electricity, lustrous and has very high melting point.

7.3 (e) Chemical Properties of Metals :

The atoms of the metals have usually 1, 2 or 3 electrons in their outermost shells. These outermost electrons are loosely held by their nuclei. Therefore, the metal atoms can easily lose their outermost electrons to from positively charged ions. For example, sodium metal can lose outermost one electron to form positively charged ions, Na^+. After losing the outermost electron, it gets stable electronic configuration of the noble gas (Ne : 2, 8), Similarly, magnesium can lose two outermost electron to from Mg^{2+} ions and aluminum can lose its three outermost electrons to from Al^{3+} ions.

$$\underset{(2,8,1)}{Na} \longrightarrow \underset{(2,8)}{Na^+} + e^-$$

$$\underset{(2,8,2)}{Mg} \longrightarrow \underset{(2,8)}{Mg^{2+}} + 2e^-$$

$$\underset{(2,8,3)}{Al} \longrightarrow \underset{(2,8)}{Al^{3+}} + 3e^-$$

The metal atoms lose electrons and form positively charged ions, therefore, the metals are called electropositive elements.

Some of the important chemical properties of metals are discussed below :

(i) Reaction with oxygen : Metals react with oxygen to from **oxides.** These oxides are **basic** in nature. For example, sodium metal reacts with oxygen of the air and form sodium oxide.

]

$$4\,Na(s) + O_2(g) \longrightarrow \underset{\text{Sodium oxide}}{2\,Na_3O(s)}$$

Sodium oxide reacts with water to form and alkali called sodium hydroxide. Therefore, sodium oxide is a basic oxide.

$$Na_2O(s) + H32O(\ell) \longrightarrow \underset{\text{Sodium hydroxide}}{2NaOH(aq)}$$

Due to the formation of sodium hydroxide (which is an alkali), the solution of sodium oxide in water turns red litmus blue (common property of all alkaline solutions).

When metal oxides are dissolved in water, they give alkaline solutions.

Similarly, magnesium is a metal and it reacts with oxygen to form magnesium oxide. However, magnesium is less reactive than sodium and therefore, heat is required for the reaction.

$$2Mg\,(s) + O_2\,(g) \xrightarrow{\text{Heat}} 2\,MgO(s)$$

Thus, when a metal combines with oxygen, it loses its valence electrons and forms positively charged metal ions. We can say that oxidation of metal takes palace.

Reactivity of metals with oxygen:

All metals **do not react** with oxygen with **equal ease.** The reactivity of oxygen depends upon the nature of the metal. Some metals **react** with oxygen even at **room temperature, some react on on heating** while still **others react** only on strong heating.

For example :

(A) Metals like sodium, potassium and calcium react with oxygen even at room temperature to form their oxides.

$4Na(s)$ +	$O_2(g)$	$\longrightarrow$	$2\ Na_2O(s)$
Sodium	Oxygen		Sodium oxide
$4K\ (s)$ +	$O_2(g)$	$\longrightarrow$	$2K_2O(s)$
Potassium	Oxygen		Potassium oxide
$2Ca(s)$ +	$O_2(g)$	$\longrightarrow$	$2\ CaO(s)$
Calcium		Oxygen	Calcium oxide

(B) Metals like magnesium and zinc do not react with oxygen at room temperature. They burn in air only on strong heating to from corresponding oxides.

$2Mg(s)$ +	$O_2(g)$	$\xrightarrow{\text{Heat}}$	$2MgO(s)$
Magnesium	Oxygen		Magnesium oxide
$2\ Zn(s)$ +	$O_2(g)$	$\xrightarrow{\text{Heat}}$	$2\ ZnO(s)$
Zinc	Oxygen		Zinc oxide

(C) Metals like iron and copper do not burn in air even on strong heating. However, they react with oxygen only on prolonged heating.

$3Fe(s)$ +	$2O_2(g)$	$\xrightarrow{\text{Heat}}$	$Fe_3O_4(s)$
Iron	Oxygen		Iron (II, III) oxide
$2\ Cu(s)$ +	$O_2(g)$	$\xrightarrow{\text{Heat}}$	$2CuO(s)$
Copper	Oxygen		Copper (II) oxide

(ii) Reaction with water : Metals react with water to form metal oxide or metals hydroxide and hydrogen. The reactivity of metals towards water depends upon the nature of the metals. Some metals react even with cold water, some react with water only on heating while there are some metals do not react even with steam. For example,

(A) Sodium and potassium metals react vigorously with cold water to form sodium hydroxide and hydrogen gas is liberated.

$2\ Na(s)$ +	$2H_2O(\ell)$	$\longrightarrow$	$2NaOH(aq)$	+	$H_2(g)$
Sodium	Cold water		Sodium hydroxide		Hydrogen
$2K\ (s)$ +	$2H_2O(\ell)$	$\longrightarrow$	$2KOH\ (aq)$	+	$H_2\ (g)$
Potassium			Potassium hydroxide		Hydrogen

The reaction between sodium and water is so violent that the hydrogen evolved catches fire.

(B) Calcium reacts with cold water to form calcium hydroxide and hydrogen gas. The reaction is **less violent.**

$Ca\ (s)$ +	$2H_2O(\ell)$	$\longrightarrow$	$Ca\ (OH)_2\ (aq) + H_2(g)$
Calcium	Cold water		calcium hydroxide

(C) Magnesium reacts very slowly with cold water but reacts rapidly with hot boiling water forming magnesium oxide and hydrogen.

$Mg\ (s)$ +	$H_2O(\ell)$	$\longrightarrow$	$MgO(s)$ +	$H_2(g)$
Magnesium	Boiling water		Magnesium oxide	

(D) Metals like zinc and aluminum react only with steam to form their corresponding oxides oxide hydrogen.

$Zn(s)$	+	$H_2O(g)$	$\longrightarrow$	$ZnO(s)$	+	$H_2(g)$
Zinc		Steam		Zinc oxide		
$2Al\ (s)$	+	$3H_2O(g)$	$\longrightarrow$	$Al_2O_3\ (s)$	+	$3H_2(g)$
Aluminum		Steam		Aluminum oxide		

(E) Iron metal does not react with water under ordinary conditions. The reactions occurs only when steam is passed over red hot iron and the products are iron (II, III) oxide and hydrogen.

$3Fe(s)$	+	$4H_2O(g)$	$\longrightarrow$	$Fe_3O_4(s)$	+	$4H_2(g)$
Iron (Red hot)		Steam		Iron (II,III) oxide		Hydrogen

(F) Metals like copper, silver and gold do not react with water even under strong conditions. The order of reactivities of different metals with water is :

Na > Mg > Zn > Fe > Cu

Reactivety with water decreases

(iii) Reaction with dilute acids : Many metals react with dilute acids and liberate hydrogen gas. Only less reactive metals such as copper, silver, gold etc. do into liberate hydrogen from dilute acids. The reactions of metals with dilute hydrochloric acid (HCl) and dilute sulphuric acid (H_2SO_4) are similar. With dil. HCl, they given metal chlorides and hydrogen whereas with dil. H_2SO_4, they give metal sulphates and hydrogen.

Dilute nitric acid (HNO_3) is an oxidising agent which oxidises metals, but does not produce hydrogen.

The reactivity of different metals is different with the same acid. For example:

(A) Sodium, magnesium and calcium react violently with dilute hydrochloric acid (HCl) or dilute sulphuric acid (H_2SO_4) liberating hydrogen gas and corresponding metal salt.

$2Na(s)$	+	$2HCl\ (aq)$	$\longrightarrow$	$2NaCl(aq)$	+	$H_2(g)$
Sodium		Hydrochloric acid		Sodium chloride		Hydrogen
$2Na(s)$	+	$H_2SO_4(aq)$	$\longrightarrow$	$Na_2SO_4(aq)$	+	$H_2(g)$
Sodium		Sulphuric acid		Sodium sulphate		Hydrogen
Similarly,						
$Mg\ (s)$	+	$2HCl\ (aq)$	$\longrightarrow$	$MgCl_2(aq)$	+	$H_2(g)$
Magnesium		Hydrochloric acid		Magnesium chloride		Hydrogen
$Mg(s)$	+	$H_2SO_4(aq)$	$\longrightarrow$	$MgSO_4\ (aq)$	+	$H_2(g)$
Magnesium		Sulphuric acid acid		Magnesium sulpahte		Hydrogen

Zn(s)	+	2HCl(aq)	$\longrightarrow$	$ZnCl_2$(aq)	+	H_2(g)
Zinc		Hydrochloric acid		Zinc chloride		Hydrogen
Zn(s)	+	H_2SO_4(aq)	$\longrightarrow$	$ZnSO_4$(aq)	+	H_2(g)
Zinc		Sulphuric acid		Zinc sulphate		Hydrogen

Similarly,

2Al (s)	+	6HCl (aq)	$\longrightarrow$	$2AlCl_3$(aq)	+	$3H_2$(g)
Aluminum		Sulphuric acid		Aluminum sulphate		Hydrogen

(B) Iron react slowly with dilute HCl or dil. H_2SO_4 and therefore, it is less reactive than zinc and aluminum.

Fe(s)	+	2HCl(aq)	$\longrightarrow$	$FeCl_2$(aq)	+	H_2(g)
Iron		Hydrochloric acid		Ferrous chloride		Hydrogen
Fe(s)	+	H_2SO_4(aq)	$\longrightarrow$	$FeSO_4$(aq)	+	H_2(g)
Iron		Sulphuric acid		Ferrous sulphate		Hydrogen

(C) Copper does not react with dill. HCl or dill H_2SO_4.

Cu(s) + HCl (aq) $\longrightarrow$ No reaction

Cu(s) + H_2SO_4(aq) $\longrightarrow$ No reaction

Therefore copper is even less reactive that iron.

The order of reactivity of different metals with dilute acid:

Na > Mg > Al > Zn > Fe > Cu

Reactivity with dill acids decreases from sodium to copper.

(iv) Reactions of metals with salt solutions: When a more reactive metals is placed in a salt solution of less reactive metal, then the more reactive metal displaces the less reactive metal from its salt solution. For example, we will take a solution of copper sulphate (blue coloured solution) and put a strip of zinc metal in the solution. It is observed that the blue colour of copper sulphate fades gradually and copper metals are deposited on the zinc strip. this means that the following reaction occurs :

Zn(s)	+	$CuSO_4$ (aq)	$\longrightarrow$	$ZnSO_4$(aq)	+	Cu(s)
Zinc		Copper sulphate (Blue solution)		Zinc sulphate (Colourless solution)		Copper

Here, zinc displaces copper from its salt solution.
However, if we take zinc sulphate solution and put a string of copper metal in this solution, no reaction occurs.

$ZnSO_4$(aq)	+	Cu(s)	$\longrightarrow$	No reaction
Zinc sulphate		Copper		

This means that **copper cannot displace zinc** metal from its solution. Thus, we can conclude that zinc is more reactive than copper. However, if we put gold or platinum strip in the copper sulphate solution, then copper is not displaced by gold or platinum. Thus, gold and platinum are less reactive than copper.

7.4 REACTIVITY SERIES OF METALS:

7.4 (a) Introduction:

We have learnt that some metals are chemically very reactive while others are less reactive or do not react at all.

On the basis of reactivity of different metals with oxygen, water acids as well as displacement reactions, the metals have been arranged in the decreasing order of their reactivities.

The arrangement of metals in order of decreasing reactivities is called reactivity series or activity series of metals.

The activity series of some common metals is given in Table. In this table, the **most reactive metal** is placed at the **top** whereas the **least reactive** metal is placed at the **bottom.** As we go down the series the chemical reactivity of metals decreases.

REACTIVITY SERIES OF METALS

त्मंबजपअपजल पदबतमेंमे	Lithium	Li	Most reactive metal
	Potassium	K	त्मंबजपअपजल कमबतमेंमे
	Barium	Ba	
डमजंसे उवतम तमंबजपअम जींद ीलकतवहमद	Sodium	Na	
	Calcium	Ca	
	Magnesium	Mg	
	Aluminum	Al	
	Zinc	Zn	
	Iron	Fe	
	Nickel	Ni	
	Tin	Sn	
	Lead	Pb	
	भ्लकतवहमद	भ्	
डमजंसे समे तमंबजपअम जींद ीलकतवहमद	Copper	Cu	
	Mercury	Hg	
	Silver	Ag	
	Gold	Au	
	Platinum	Pt	Least reactive metal

7.4 (b) Reasons for Different Reactivities:

In the activity series of metals, the basis of reactivity is the tendency of metals to lose electrons. If a metals can lose electrons easily to form positive ions, it will react readily with other substances. Therefore, it will be a reactive metal. On the other hand, if a metal loses electrons less rapidly to form a positive ion, it will react slowly with the other substances. Therefore, such a metal will be less reactive. For example, alkali metals such as sodium and potassium lose electrons very readily to from alkali metal ions, therefore, they are very reactive.

7.4 (c) Displacement of Hydrogen from Acids by Metals :

All metals above hydrogen in the reactivity series (i.e. more active than hydrogen) like zinc, magnesium, nickel can liberate hydrogen from acids like HCl and H_2SO_4. These metals have greater tendency to lose electrons than hydrogen. Therefore, the H^+ ions in the acids will accept electrons and give hydrogen gas as :

$$\underset{\text{Metals}}{M} \longrightarrow M^{+}(aQ) + e^{-}$$

$$\underset{\text{(From acid)}}{H^{+}(aq) + e^{-}} \longrightarrow H$$

$$H + H \longrightarrow H_2\uparrow$$

The metals which are below hydrogen in the reactivity series (i.e. les reactive than hydrogen) like copper, silver, gold cannot liberate hydrogen form acids like HCI, H_2SO_4 etc. These metals have lesser tendency to lose electrons than hydrogen. Therefore, they cannot lose electrons to H^+ ions.

7.4 (d) Reactivity Series and Displacement Reactions :

The reactivity series can also explain displacement reactions. In general, a more **reactive metal** (placed higher in the activity series) can displace the less reactive metal from its solution. For example, zinc, displaces copper form its solution.

$Zn\,(s) + CuSO_4\,(aq) \longrightarrow ZnSO_4\,(aq) + Cu(s)$

7.4 (e) Usefulness of Activity Series:

The activity series is very useful and it gives the following information:

(i) The metal which is higher in the activity series is more reactive than the other. Lithium is the most reactive and platinum is the least reactive.

(ii) The metals which have been placed above hydrogen are more reactive than hydrogen and these can displace hydrogen from its compounds like water and acids to liberate hydrogen gas.

(iii) The metals which are placed below hydrogen are less reactive than hydrogen and these cannot displace hydrogen from its compounds like water and acids.

(iv) A more reactive metal (placed higher in the activity series) can displace the less reactive metal from its solution.

(v) Metals at the top of the series are very reactive and, therefore, they do not occur free in nature. The metals at the bottom of the series are least reactive and, therefore, they normally occur free in nature. For example, gold, present in the reactivity series is found in Free State in nature.

DAILY PRACTIVE PROBLESM # 7

OBJECTIVE DPP -7.1

1. Which of the following properties is not a characteristic of metals ?
   (A) Metallic lusture (B) High density
   (C) Hardness (D) Low melting and boiling point

2. Which of the following metals generally occur in liquid sate ?
   (A) Mercury (B) Bromine (C) Gallium (D) A & C both

3. Reactivity of zinc is______ than hydrogen.
   (A) less (B) more
   (C) equal (D) sometimes more sometimes less

4. $Zn + xHCl \longrightarrow ZnCl_2 + Z$,
In above equation A & x are
(A) H_2, 2 (B) Cl_2, 1 (C) H_2, 3 (D) H_2, 4

5. When sodium reacts with cold water, then the product formed will be-
(A) Na_2O (B) NaOH (C) Na_2CO_3 (D) All of these

6. What is the decreasing order of reactivity of following metals ?
Na, Al, Cu, Ag, Fe
(A) Na > K > Al > Cu > Ag > Fe (B) K > Na > Al > Cu > Fe > Ag
(C) K > Na > Al > Fe > Cu > Ag (D) K > Na > Al > Fe > Ag > Cu

7. When a metal is added to dilute HCl solution, there is no evolution of gas. Metals if -
(A) K (B) Na (C) Ag (D) Zn

8. On addition of which metal, copper sulphate solution (Blue colour) will be changed to colourless solution >
(A) Fe (B) Ag (C) Zn (D) Hg

9. $Zn + H_2O$ (Steam) $\longrightarrow$ A + B
In the above equation (A) and (B) are
(A) Zn & H_3 (B) ZnH_2 & O_2 (C) ZnO_2 & O_2 (D) ZnO & H_2

10. Which of the following metals reacts vigorously with oxygen?
(A) Zinc (B) Magnesium (C) Sodium (D) Copper

SUBJECTVE DPP - 7.2

1. Describe the physical properties of metals >

2. Write the chemical equation of chemical reaction of zinc metal with the following -
(a) H_2SO_4 (b) H_2O (c) O_2

3. What is an activity series of metals ? Arrange the metals Zn, Mg. Al, Cu and Fe in the decreasing order of reactivity.

4. What would you observe when you put :
(i) some zinc pieces in the blue copper sulphate solution ?
(ii) some copper pieces in green ferrous sulphate solution ?

5. Name two metals which occur in nature in the free state.

6. Identify the most reactive and least reactive metal from the following -
Hg, Na, Fe, Ag.

7. Name a gas which is always produced when a reactive metal reacts with a dulute acid.
Write a chemical equation supporting your answer.

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➤METALS AND NON-METALS◀

8.1 HOW METALS REACT WITH NON-METALS :

Octet Rule : Octet rule was given by **G.N. Lewis and W.Kossel in 1916.**
According to actet rule "an atom whose outermost shell contains 8 electrons (octet) is stable."

This rule, however, does not hold good in case of certain small atoms like helium (He) in which presence of 2 electrons (duplet) in the outermost shell in considered to be the condition of stability.

Examples of elements whose atoms have fully filled or 8 e^- in their outermost shell are –

Element	Symbol	Atomic Number	Electronic configuration	No. of valence electrons
Neon	Ne	10	2,8	8
Argon	Ar	18	2,8,8	8
Krypton	Kr	36	2,8,18,8	8

All noble gases contain 8 valence electrons (except He in which 2 valence electrons are present) and are stable. They do not usually form bonds with other elements.

Atoms combine with one another to achieve the inert gas electron arrangement and become stable. Atoms from chemical bonds to achieve stability by acquiring the inert gas configuration or by completing their octet or duplet (in case of small atoms) in outermost shell. An atom can achieve the inert gas electron arrangement in three ways -

(i) by losing one or more electrons .

(ii) by gaining one or more electrons.

(iii) by sharing one or more electrons.

Store in your memory

Noble gases do not usually from bonds with other elements. because they are stable. So. atoms of elements have the tendency to combine with one another to achieve the inert gas configuration.

8.2 CONCEPT OF IONIC BOND :

Except the elements of group 18 of the periodic table all the elements for the remaining group, at normal temperature and pressure, are not stable in independent state. These elements from stable compounds either by combining with the other atoms or with their own atoms. When in gross electronic configuration of the elements there are 8 electrons present then these elements do not take part in the chemical reaction because atoms containing 8 electrons in their outermost shell are associated with extra stability and less energy.

Atoms with other electronic configuration, which do not contain eight electrons in their outermost shell, are unstable and to achieve the stability they chemically combine in such a manner that they achieve eight electrons in their outermost shell.

Two or more than two types of atoms mutually combine with each other to achieve stable configuration of eight valence electrons. Attempt to achieve eight electrons in the outermost orbit of a element is the reason behind its chemical reactivity or chemical bonding.

8.3 IONIC OR ELECTROVALENT BOND:

This bond is formed by the atoms of electropositive and electronegative elements. Electropositive elements lose electrons in chemical reaction and electronegative elements gain electrons in chemical reaction. When an atom of electropositive element come in contact with that of an electronegative element then the electropositive atom loses electron & becomes positively charged, while the electronegative atom gains the electron to become negatively charged. Electrostatic force of attractions works between the positively and negatively charged ions due to which both ions are bonded with each other. As a result, a chemical bond is produced between the ions, which is known as Ionic or Electrovalent compound.

Number of electrons donated or accepted by any element is called Electrovalence.

In an ionic compound every cation is surrounded by a fixed number of anions and every **anion is** surrounded by a fixed number of cations and they are bonded in a **fixed geometry** in a three dimensional structure.

Example : Sodium chloride compound.

Sodium atom (Electropositive element) by losing an electron from its outermost orbit, gets converted into a cation and attains noble gas like stable configuration.

Energy required for this process is called “ionization potential.”

$$\overset{x}{Na} + IE \longrightarrow Na^{+} + e^{-}$$

(2, 8, 1) (2, 8)

Chlorine atom (Electronegative element) accepts the electron donated by sodium atom in its outermost orbit and forms chloride anion.
In this process energy is released which is known as “electron affinity.”

$$:\ddot{\underset{..}{Cl}}: + e^- \longrightarrow :\ddot{\underset{..}{Cl}}:^- + EA$$

(2, 8, 7) (2, 8, 8)

Due to the opposite charges on the Na^+ and Cl^- ions, they are bonded by electrostatic force of attraction to from NaCl compound.

$$Na^+ + Cl^- \longrightarrow [Na]^+ [Cl]^- \text{ or } NaCl$$

Here electrovalent of sodium and chlorine atom is one.

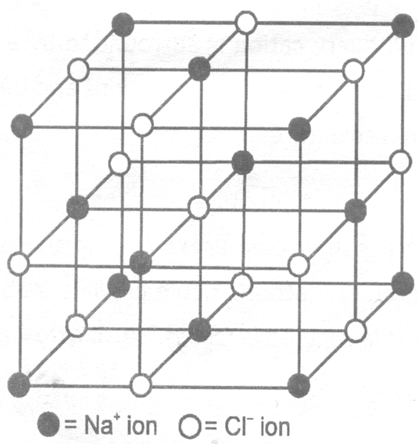
For the formation of ionic bond, it is necessary that the ionization potential of electropositive element should be less and the electron affinity of electronegative element should be high.

8.3 (a) Properties of Ionic Compounds:

(i) Ionic compounds consist of ions: All ionic compounds consist of positively and negatively charged ions and not molecules. For example, sodium chloride consists of Na^+ and Cl^- ions, magnesium fluoride consists of Mg^{2+} and F^- ions and so on.

(ii) Physical nature : Ionic compounds are solid and relatively hard due to strong electrostatic force of attraction between the ions of ionic compound.

(iii) Crystal structure : X-ray studies have shown that ionic compounds do not exist as simple single molecules as Na^+Cl^-, This is due to the fact that the forces of attraction are not restricted to single unit such as Na^+ and Cl^- but due to uniform electric field around and ion, each ion is attracted to a large number of other ions. For example, one Na^+ ion will not attract only one Cl^- ion but it can attract as many negative charges as it an. Similarly, the Cl^- ion will attract several Na^- ions. As a result, there is a regular arrangement of these ions in three dimensions as shown in diagram. Such a regular arrangements is called crystal lattice.

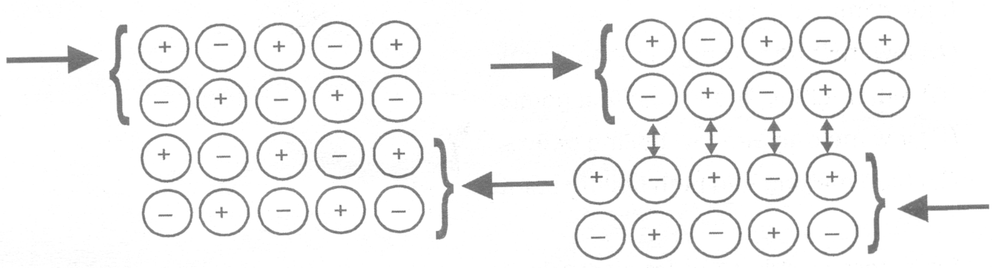


Lattice structure of Sodium chloride

(iv) Melting point and boiling point : Strong electrostatic force of attraction if present between ions of opposite charges. To break the crystal lattice more energy is required so their melting points and boiling points are high.

(v) Solubility : Ionic compounds are generally soluble in polar solvents like water and insoluble in no - polar solvents like carbon tetrachloride, benzene, ether alcohol etc.

(vi) Brittle nature: Ionic compounds on applying external force or pressure are broken into small pieces, such substances are known as brittle and this property is known as brittleness. When external force is applied on the ionic compound, layers of ions slide over one another and particles of the same charge come near to each other as a result due to the strong repulsion force, crystals of compounds are broken.



BRITTLE NATURE OF IONIC COMPOUNDS

(vii) Electrical conductivity : Electrical conductivity in any substance is due to the movement of free electrons of ions. In metals electrical conductivity is due to the free movement of valency electrons. As ionic compound exhibits electrical conductivity due to the movement of ions either in the fused state or in the soluble state in the polar solvent. But in the solid state due to strong electrostatic force of attraction free ions are absent so they are insulator in the solid state.

DAILY PRACTIVE PROBLEMS # 8

OBJECTIVE DPP - 8.1

1. Octet rule was given by -
(A) Rutherford (B) Soddy (C) Lewis & Kossel (D) None of these

2. Exception of octet rule is -
(A) K (B) Ca (C) N (D) He

3. Ionic bond is formed by -
(A) loss of electrons only. (B) gain of electrons only.
(C) loss and gain of electrons both. (D) sharing of electrons.

4. Ionic bond is formed between -

(A) two electropositive elements.

(B) two electronegative elements.

(C) Electropositive & electronegative elements.

(D) None of these

5. During formation of ionic bond -

(A) there is force of repulsion between two negative ions.

(B) there is force of repulsion between two positive ions.

(C) there is force of attraction between positive & negative ions.

(D) none of these.

6. In the formation of ionic bond, cation is formed by-
(A) gain of electron (s). (B) loss of electron(s).
(C) sharing of electron(s). (D) None of these

7. Ionic compound have -

(A) low melting and high boiling points.

(B) high melting and low boiling points.

(C) low melting and low boiling points.

(D) high melting and high boiling points.

8. Ionic compounds conduct electricity in-
(A) solid state (B) fused state.
(C) gaseous state. (D) Do not conduct electricity at all.

9. Ionic compounds are soluble in-
(A) water (B) benzene (C) ether (D) alcohol

10. Physical nature of most of the ionic compounds is-
(A) solid (B) liquid (C) gas (D) May exist in any state.

SUBJECTIVE DPP - 8.2

1. Define octet rule.

2. Define electrovalency.

3. Explain the brittle nature of ionic compounds.

4. Why ionic compounds have high melting and boiling points ?

5. Why ionic compounds show electrical conductivity is fused or soluble state ?

Teko CLASSES ......the support

➤METALS AND NON-METALS⮜

9.1 OCCURRENCE OF METALS :

All metals are present in the earth's crust either in the free state or in the form of their compounds. Aluminum is the most abundant metal in the earth's crust. The second most abundant metal is iron and thier one is calcium.

9.1 (a) Native and Combined State of Metals :

Metals occur in the crust of earth in the following two states -

(i) Native state of free state: A metal is said to occur in a free or a native state when it is found in the crust of the earth in the elementary or uncombined form.
The metals which are very uncreative (lying at the bottom of activity series) are found in the free stae. These have no tendency to react with oxygen and are not attacked by moisture, carbon dioxide of air or other no-metals. Silver, copper, gold and platinum are some examples of such metals.

(ii) Combines state : A metal is said to occur in a combined state if it is found in nature in the form of its compounds. e.g. Sodium, magnesium etc. Copper and silver are metals which occur in the free state as well as in the combined state.

9.2 MINERALS AND ORES :

The natural substances in which metals or their compounds occur either in native state or combined state are called minerals.
The minerals are not pure and contain different types of other impurities. The impurities associated with minerals are collectively known as gangue or matrix.
The mineral from which the metal can be conveniently and profitably extracted, is called an ore.
For example, aluminum occurs in the earth's crust in the form of two minerals, bauxite (Al_2,O_3. $2H_2O$) and caly (Al_2O_3. $2SiO_2$. $2H_2O$). Out of these two, aluminum can be conveniently and profitably extracted from bauxite. So, bauxite is an are of aluminum.

Oxygen is the most abundant element on earth's crust.

9.2 (a) Types of Ores :

The most common ores of metals are oxides, sulphides, carbonates, sulphates, halides, etc. In general, very uncreative metals (such as gold, silver, platinum etc.) occur in elemental form or Free State.

(i) Metals which are only slightly reactive occur as sulphides (e.g., CuS, Pbs etc.).

(ii) Reactive metals occur as **oxides** (e.g., MnO_2, AO_2O_3 etc.)

(iii) Most reactive metals occur as salts as **carbonates, subparts, halides** etc.

SOME COMMON ORES ARE LISTED IN THE TABLE

Nature	Metal	Name of the ore	Composition
Oxide ores	**Aluminum**	**Bauxite**	$Al_2O_3.2H_2O$
	Copper	**Cuprite**	Cu_2O
	Iron	**Magnetite**	Fe_3O_4
		Heamatite	Fe_2O_3
Sulphide ores	**Copper**	**Copper pyrites**	$CuFeS_2$
		Copper glance	Cu_2S
	Zinc	**Zinc blende**	**Zns**
	Lead	**Galena**	**PbS**
	Mercury	**Cinnabar**	**HgS**
Carbonate ores	**Calcium**	**Limestone**	$CaCO_3$
	Zinc	**Calamine**	$ZnCO_3$
Halide ores	**Sodium**	**Rock salt**	**NaCl**
	Magnesium	**Carnallite**	$KCl.\ MgCl_2.\ 6H_2O$
	Calcium	**Fluors par**	CaF_2
	Silver	**Horn silver**	**AgCl**
Sulphate ores	**Calcium**	**Gypsum**	$CSO_4\ 2H_2O$
	Magnesium	**Epsom salt**	$MgSO_4.7H_2O$
	Barium	**Barytes**	$BaSO_4$
	Lead	**Anglesite**	$PbSO_4$

9.3 METALLURGY:

The process of extracting pure metals from their ores and then refining them for use is called metallurgy.
In other words, the process of metallurgy involves extraction of metals from their ores and then refining them from use. The ores generally contain unwanted impurities such as sand, stone, earthy particles, limestone, mica, etc., these are called gangue or matrix.

The process of metallurgy depends upon the nature of the ore, nature of the metals and they types of impurities present. Therefore, there is not a single method for the extraction of all metals. However, most of the metals can e extracted by a general procedure which involves the following steps.

Various steps involved in metallurgical processes are -

(a) Crushing and grinding of the ore.

(b) Concentration of the ore or enrichment of the ore.

(c) Extraction of metal from the concentrated ore.

(d) Refining or purification of the impure metal.

These steps are briefly discussed below -

9.3 (a) Crushing and Grinding of Ore :

Most of the ores occur as big rocks in nature. They are broken into small pieces with the help of crusher. These pieces are then reduced to fine powder with the help of a ball mill or a stamp mill.

9.3 (b) Concentration of Ore or Enrichment of Ore :

The process of removal of unwanted impurities (gangue) from the ore is called ore concentration or ore enrichment.

(i) Hydraulic washing (washing with water) :

Principle: This method is based upon the difference in the densities of the ore particles and the impurities (gangue).
Ores of iron, tin and lead are very heady and, therefore, they are concentrated by this method.

(ii) Front floatation process:

Principle: this method is based on the principle of difference in the wetting properties of the ore and gangue particles with water and oil.
This method is commonly used for sulphide ores.

(iii) Magnetic separation :

Principle: This method depends upon the difference in the magnetic properties of the ores and gangue.
This method is used for the concentration of **heamatite**, an ore of iron.


Store in your memory

The froth floatation process if commonly used for the sulphide ores copper, zinc, lead etc.

DIALY PRACTICE PROBLEMS # 9

OBJECTIVE DPP - 9.1

1. Which of the following is/are oxide ore(s) ?
(A) Bauxite (B) Cuprite (C) Haematite (D) All of these

2. Horn silver is a/an -
(A) sulphate ore (B) halide ore (C) sulphide ore (D) oxide ore

3. Carnallite is -
(A) KCl, $MgCl_2$ (B) KCl. $MgCl_2$, $3H_2O$ (C) KCl. $MgCl_2$. $6H_2O$ (D) KCl, $MgCl_2$, H_2O

4. Match column A with column B and select the correct option -

Column A	**Column B**
(Ore)	(Nature of ore)
(a) Copper glance	(i) Sulpahte ore
(b) Calamine	(ii) Halide ore
(c) Rock salt	(iii) Sulphide ore
(d) Epsom salt	(iv) Carbonate ore

(A) a(i), b(ii), c(iii), d(iv) (B) a(iv), b(ii), c(iii), d(i)
(C) a(iii), b(iv), c(ii), d(i) (D) a(iv), b(i), c(ii), d(iii)

5. Removal of impurities from ore is known as -
(A) crushing and grinding (B) concentration of ore
(C) minerals (D) gangue

6. Which of the following methods is used in the concentration of haematite ore ?
(A) Hydraulic washing (B) Magnetic separation
(C) Froth floatation process (D) None of these

7. Forth floatation method is used for the concentration of -
(A) oxide ores (B) sulphide ores (C) sulphate ores (D) halide ores

8. Which of the following methods is based on the principle of the difference in the wetting properties of the ore and gangue particles with water and oil ?
(A) Magnetic separation (B) Front floatation process
(C) Hydraulic washing (D) None of these

9. Which of the following is most abundant metal on the earth's crust ?
(A) Iron (B) Aluminum (C) Calcium (D) Oxygen

10. Which of the following metal is found in native state ?
(A) Sodium (B) Zinc (C) Gold (D) Iron

SUBJECTIVE DPP- 9.2

1. Explain the difference between ores and minerals ?

2. Comment on native and combined states of metals.

3. What is gangue ?

4. Which process is used for the enrichment of
(i) sulphide ores (ii) oxide ore

5. Give chemical compositions of the following ores -
(i) bauxite (ii) gypsum
(iii) galena (iv) rock salt

Teko CLASSES ....the support

➤METALS AND NON-METALS◄

10.1 EXTRACTION OF THE METAL FROM THE CONCENTRATED ORE :

The metal is extracted from the concentrated ore by the following steps :

(a) Conversion of the concentrated ore into its oxide : The production of metal from the concentrated ore mainly involves reduction process. This can be usually done by two processes known as calcination and roasting process. The method depends upon the nature of the ore.

(b) Conversion of oxide to metal y reduction process

10.1 (a) Conversion of Ore into Metal Oxide :

These are briefly discussed below :

(i) Calcination : It is the process of heating the concentrated ore in the absence of air.

The calcination process is used for the following changes :

- to convert carbonate ores into metal oxide.
- to remove water from the hydrated ores.
- to remove volatile impurities from the ore.

For example

$ZnCO_3(s) \xrightarrow{\text{Calcination}} ZnO(s) + CO_2(g)$

Zinc carbonate — Zinc oxide — carbon dioxide

$FeCO_3(s) \xrightarrow{\text{Calcination}} FeO(s) + CO_2(g)$

Siderite — Iron (I) oxide — Carbon dioxide

(ii) Roasting : It is the process of heating the concentrated ore strongly in the presence of excess air. This process is used for converting sulphide ores to metal oxide. In this process, the following changes take place :

- the sulphide ores undergo oxidation to their oxides.
- moisture is removed
- volatile impurities are removed.

For example :

$2ZnS + 3O_2 \xrightarrow{\text{Roasting}} 2ZnO(s) + 2SO_2(g)$

Zing (Zinc blende ore) — Oxygen (From air) — Zinc oxide — Sulphur

$4FeS_2(s) + 11O_2(g) \xrightarrow{\text{Roasting}} 2Fe_2O_3(s) + 8SO_2(g)$

Iron pyrites — Oxygen — Ferric oxide — Sulphur Dioxide

Calcination is used for hydrated and carbonate ores and roasting is used for sulphide ores.

10.1 (b) Conversion of Metal Oxide to Metal:

The metal oxide formed after calcination or roasting is converted into metal by reduction. The method used for reduction of metal oxide depends upon the nature and chemical reactivity of metal.
The metals can be grouped into the following three categories on the basis for their reactivity:

- Metals of low reactivity.
- Metals of medium reactivity.
- Metals of high reactivity.

These different categories of metals are extracted by different technique. the different steps involved in separation are as follows :

(i) Reduction by heating : Metals placed low in the reactivity series are very less reactive. They can be obtained from their oxides by simple heating in air.

$2HgS(s)$	+	$3O_2(g) \xrightarrow{\text{Roasting}}$	$2HgO(s)$	+	$2SO_2(g)$
Mercuric sulphide		Oxygen	Mercuric oxide		Sulphur dioxide

$2HgO(s)$	$\xrightarrow{\text{Roasting}}$	$2Hg(\ell)$	+	$O_2(g)$
Mercuric oxide		Mercury metal		Oxygen

(ii) Chemical Reduction (For metals in the middle of the reactivity series):
The metals n the middle of the reactivity series, such as iron, zinc, lead, copper etc. are moderately reactive. These are usually present as sulphides or carbonates. Therefore, before reduction the metal sulphides and carbonates must be converted to oxides. This is done by roasting and calcination. The oxides of these metals cannot be reduced by heating alone. Therefore, these metal oxides are reduced to free metal by using chemical agents like carbon, aluminum, sodium or calcium.

(A) Reduction with carbon : The oxides of moderately reactive metals (occurring in the meddle of reactivity series) like zinc, copper, nickel, tin, lead etc. can be reduced by using carbon as reducing agent.

$AnO(s)$	+	$C(s)$	$\xrightarrow{\text{Heat}}$	$Zn(s)$	+	$CO(g)$
Zinc oxide		Carbon (Reducing agent)		Zinc metal		Carbon monoxide

$Fe_2O_3(s)$	+	$3C(s)$	$\longrightarrow$	$2Fe(s)$ +	$3CO(g)$
Ferric oxide		Carbon		Iron Metal	Carbon monoxide

$PbO(s)$	+	$C(s)$	$\longrightarrow$	$Pb(s)$ +	$CO(g)$
Lead oxide		Carbon		Lead metal	Carbon monoxide

One disadvantage of using carbon as reducing agent is that small traces of carbon are added to metal as impurity. Therefore, it contaminates the metals.

Coke is very commonly used as a reducing agent because it is cheap.

(B) Reduction with carbon monoxide: Metals can be obtained from oxides by reduction with carbon monoxide in the furnace.

$Fe_2O_3(s)$	+	$3CO(g)$	$\xrightarrow{\text{Heat}}$	$2Fe(s)$ +	$3CO_2(g)$
Ferric oxide		Carbon monoxide		Iron	Carbon dioxide

(C) Reduction with aluminum : Certain metal oxides are reduced by aluminum to metals.

$3MnO_2$ (s)	+	$4Al(s)$	$\xrightarrow{\text{Heat}}$	$3Mn(s)$ +	$2Al_2O_3(s)$
Manganese dioxide		Aluminum		Manganese	Aluminum oxide

$Cr_2O_3(s)$	+	$2Al$ (s)	$\xrightarrow{\text{Heat}}$	$2Cr$ (s) +	Al_2O_3 (s)
Chromium oxide		Aluminum		Chromium	Aluminum oxide

Reduction of metals oxides with aluminum is known as aluminothermy or thermite process.

(iii) Reduction of electrolysis or electrolytic reduction : The oxide of **active metals** (which are high up in the activity series) are very **stable** and cannot be reduced by **carbon or aluminum.** These metals are commonly extracted by the **electrolysis** of their fused salts using suitable electrodes. This is also called **electrolytic reduction i.e.** reduction by electrolysis.
For example, aluminum oxide is very stable and aluminum cannot be prepared by reduction with carbon. It is prepared by the electrolysis of molten alumina (Al_2O_3).

Al^{3+}	+	$3e^-$	$\xrightarrow{\text{Heat}}$	Al
Aluminum ion (From molten alumina)		Electron (From cathode)		Aluminum (At cathode)

It may be noted that during electrolytic reduction of molten salts, the metals are always obtained at the cathode (negative electrode).


Store in your memory

The process of extraction of metals by electrolysis process is called electrometallurgy.

10.2 PURIFICATION OR REFINING OF METALS :

The metal obtained any of the above methods is usually impure and is known as **crude** metal. The process of purifying the crude metal is called **refining.**

10.2 (a) Liquation :

This method is use for refining the metals having low melting points, such as tin, lead, bismuth etc. This is based on the principle that the metal to be refined is easily fusible (melt easily (but the impurities do not fuse easily.

10.2 (b) Distillation:

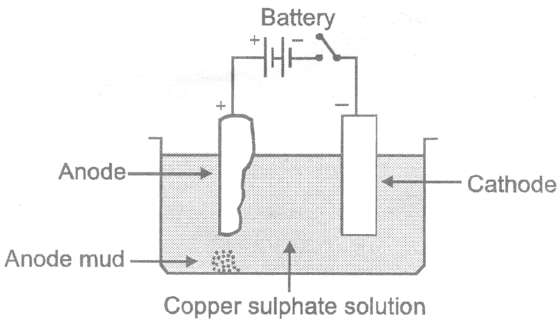
This method is used for the **purification of volatile metals** (which form vapours readily) such as mercury and zinc.

10.2 (c) Electrolytic Refining :

This is most general and widely used method for the refining of impure metals. Many metals such as copper, zinc, tin, nickel, silver, gold etc. are refined electrolytically. It is based upon the phenomenon of electrolysis. In this method, the crude metal is cast into thick rods and are made as anodes, while the thin sheets of pure metal are made as cathodes, An aqueous solution of some salt of the metal is used as an electrolyte. On passing current through the electrolyte, the pure metal from the anode dissolves into the electrolyte. An equivalent amount of pure metal from the electrolyte is deposited on the cathode. The soluble impurities go in the solution whereas the insoluble impurities settle down at the bottom of the node and are known as anode mud. In this way, the pure metal from anode goes into electrolyte and from electrolyte it goes to the cathode.

At anode : $\underset{\text{Copper (from impure anode)}}{Cu} \xrightarrow{\text{Oxidation}} Cu^{+2} + 2e^-$

At cathode : $\underset{\text{Copper ion}}{Cu^{2+} + 2e^-} \xrightarrow{\text{Reduction}} \underset{\text{Copper (deposited at cathode)}}{CU}$



In electrolytic refining impure metal is made anode and pure metal is made cathode.

Zone refining and Van Arkel method are used for obtaining metals (Si, Ge etc.) of very high purity for certain specific applications.

DAILY PRACTICE PROBLESM # 10

OBJECTIVE DPP - 10.1

1. Heating of concentrated ore in absence of air for conversion in oxide ore in known as -
(A) roasting (B) calcination (C) reduction (D) none of these

2. Process of roasting and calcination takes place in-
(A) bessemer converter. (B) blast furnace.
(C) reverberatory furnace. (D) electrolytic cell.

3. Which reducing agent is used in chemical reduction ?
(A) C (B) CO (C) Al (D) All of these

4. Which of the following is used in reduction of alumina ?
(A) Coke (B) Carbon monoxide (C) Aluminum (D) Electricity

5. For purification of which metal, liquation method is used ?
(A) Tin (B) Lead (C) Bismuth (D) All of these

6. Which method is used in purification of mercury ?
(A) Liquation (B) Distillation (C) Electrolytic refining (D) Chemical reduction

7. Which of the following methods is used for obtaining metals of very high purity ?
(A) Distillation (B) Zone refining (C) Liquation (D) Electrolytic refining

8. Which of the following methods is not used in purification of metals ?
(A) Calcination (B)Liquation (C) Distillation (D) None of these

9. Anode mud is obtained in which process?
(A) Roasting (B) Zone refining (C) Electrolytic refining (D) Calcination

10. In thermite process reducing agent is -
(A) C (B) Cl (C) Al (D) None of these

SUBJECTIVE DOO - 10.2

1. Describe methods of extraction of the metal from the concentrated ore ?

2. Define calcination.

3. Name a method for obtaining metals of very high purity.

4. Which method is used for refining of volatile metals ?

5. Name the products obtained when -
(i) zinc sulphide is roasted (ii) lime stone is calcined

➤METALS AND NON-METALS⮜

11.1 CORROSION OF METALS :

Surface of many metals is easily attacked when exposed to atmosphere. The react with **air or water** present in the environment and form **undesirable compounds** on their surface. These undesirable compounds are generally **oxides.**
Thus, corrosion is a process of **deterioration** of metal as a result of its reaction with air or water (present in environment) surrounding it.

11.1 (a) Corrosion of Iron:

iron corrodes readily when exposed to **moisture** and gets covered with a **brown flaky** substance called **rust.** This is also called Rusting of Iron. Chemically, the rust if **hydrated iron (III) oxide,** $Fe_2O_3.XH_2O$. Rusting is an **oxidation process** in which iron metal is slowly oxidized by the action of air (in presence of water). Therefore, rusting of iron takes place under the following conditions :

- Presence of air (or oxygen)
- Presence of water (moisture)

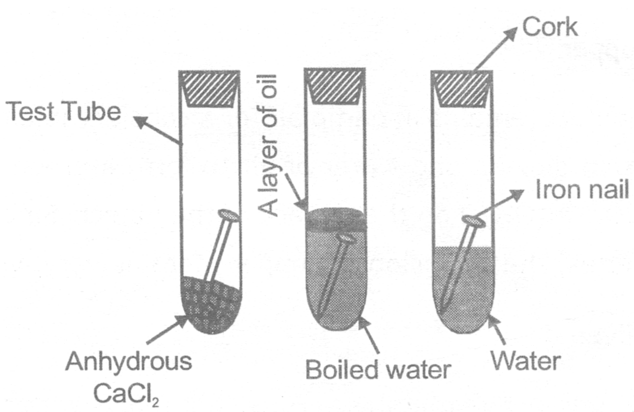
More the reactivity of the metal, the more will be the possibility of the metal getting corroded.

(i) Experiment to show that rusting of iron requires both air and water -
We take three test tubes and put one clean **iron nail** in each of the three test tubes :

(A) In the first test tube containing **iron nail,** we put some **anhydrous calcium chloride** to absorb water (or moisture) from the damp air present in the test tube and make it dry.

(B) In the second test tube containing iron nail, we put **boiled water** because boiled water does not contain any dissolved air or oxygen in it. **A layer of oil** is put over boiled water in the test tube to prevent the outside air from mixing with boiled water.

(C) In the third test tube containing an iron nail, we put **unboiled water** so that about two-third of the nail is immersed in water and the rest is above water exposed to damp air.
After one week, we observe the iron nails kept in all the three test tubes.



Rusting of iron

(ii) We will obtain the following observations from the experiment :

(A) No rust in seen on the surface of iron nail kept in dry air in the first test tube. This tells us that **rusting of iron** does not takes place in **air alone.**

(B) No rust is seen on the surface of iron nail kept in air free boiled water in the second test tube, This tells us that **rusting of iron** does not take place in **water alone.**

(C) Red brown rust is seen on the surface of iron nail kept in the presence of the air and water in the third test tube. This tells us that **rusting of iron** takes place in the presence of **both air and water** together.

(iii) Prevention of rusting

(A) Corrosion of metals can be prevented by coating the metal surface with a thin layer of **pant, varnish or grease.**

(B) Iron is protected from rusting by coating it with a thin layer of another metal which is more reactive that iron. This prevents the loss of electrons from iron because the active metal loses electrons in process of covering iron with **zinc** is called **galvanization.** Iron is also coated with other metals such as **tin** known as **tin coating.**

(C) By alloying: Some metals when alloyed with other metals become more resistant to corrosion. For example, when iron is alloyed with chromium and nickel, it form stainless steel. This is resistant to corrosion and does not rust at all.

(D) To decrease rusting of iron, certain **antirust solutions** are used. For example, solutions of **alkaline phosphates** are used as antirust solutions.

11.1 (b) Corrosion of Aluminum :

Due to the formation of a dull layer of aluminum oxide when exposed to moist air, the aluminum metal loses its shine very soon after use. This **aluminum oxide layer** is very tough and prevents the metal underneath from further corrosion (because moist air is not able to pass through this aluminum oxide layer). This means sometimes corrosion is useful.

11.1 (c) Corrosion of Copper

When a copper object remains in damp air for a considerable time, then copper reacts slowly with carbon dioxide and water of air to form a green coating of basic copper carbonate [$CuCO_3$, Cu(OH($_2$] on the surface of the object. Since copper metal is low in the reactivity series, the corrosion of copper metal is very, very slow.

11.1 (d) Corrosion of Silver :

Silver is a highly uncreative metal, so it does not reacts with oxygen of air easily. But, air usually contains a little of sulphur compounds such as hydrogen sulphide gas (H_2S), which react slowly with silver to form a **black coating of silver sulphide** (Ag_2S). Silver ornaments gradually turn black due to the formation of a thin silver sulphide layer on their surface and silver is said to be tarnished.

11.2 ALLOYS :

An alloy is a homogenous mixture of two or more metals or a metal and a non-metal. For example, iron is the most widely used metal. But it is never used in the pure form. this is because iron is very soft and stretches easily when not. But when it is mixed with a small amount of **carbon** (about 0.05%), it becomes **hard and strong**. The new form of iron is called **steel.**

11.2 (a) Objective of Ally Making :

Alloys are generally prepared to have certain specified properties which are not possessed by the contituent metals. The main objects of ally-making are :

(i) To increase resistance to corrosion : For example, stainless steel is prepared which has more resistant to corrosion than iron.

(ii) To modify chemical reactivity : The chemical reactivity of sodium is decreased by making an alloy with mercury which is known as sodium amalgam.

(iii) To increase the hardness: Steel, an alloy of iron and carbon is harder than iron.

(iv) To increase tensile strength: Magnesium is an alloy or magnesium and aluminum.
It has greater tensile strength as compared to magnesium and aluminum.

(v) To produce good casting: Type metal is an alloy of lead, tin and mercury.

(vi) To lower the melting point: For example, solder is an alloy of lead and tin (50) Pb and 50% Sn). It has a low melting point and is used for welding electrical wires together.

11.2 (b) Some Important Alloys :

The approximate composition and used of some important alloys are given below :
(i) Steel : Steel is an alloy of **iron and carbon** containing **0.1 to 1.5% carbon.** Steel is very **hard, tough and strong.** It is used for making rails, screws, girders, bridges, railway lines etc. Steel can also be used for the contraction for building, vehicles, ships etc.

(ii) Allow Steels : Steel obtained by the addition of some other elements such as chromium, vanadium, titanium, molybdenum, manganese, cobalt or nickel to carbon steel are called Ally Steel.

(iii) Alloys of Aluminum : This common alleys of aluminum are :

(A) Duralumin. It is an alloy containing aluminum, copper and traces of magnesium and manganese. Its percentage composition is - AO\I 95%, Cu = 4%, Mg = 0.5 % Mn = 0.5 % It is **stronger** than pure aluminum, Since duralumin is **light** and yet **strong,** it is used for making bodies of aircrafts, helicopter, jets and kitchenware's like pressure cookers etc.

(B) Magnesium. It is an alloy of **aluminum and magnesium** having the composition: Al - 95%, Mg = 5% It is very light and hard. It is more hard than pure aluminum. It is used for making light instruments, balance beams, pressure cookers etc.

(C) Alnico . It is an alloy containing aluminum, iron nickel, and cobalt. It is highly magnetic in nature and can be used for making powerful magnets.

(iv) Alloys of Copper: The important alloys of copper are Brass and Bronze.

(A) Brass - It is an alloy of copper and zinc having the composition = Cu = 80% Zn = 20% Brass is more malleable and more strong than pure copper. It is used for making cooking utensils, condenser sheets, pipes, hardware, nuts, bolts, screws, springs etc.

(B) Bronze - It is an alloy of copper and tin having the composition : Cu = 90% Sn = 10% Bronze is very though and highly resistant to corrosion. It is used for making utensils, statues, cooling pipes, coins, hardware etc.

(C) German Silver - It is an alloy of copper, zinc and nickel having the composition: Cu = 60%, Zn = 20%, Ni = 20%. It is used for making silverware, utensils and for electroplating.

(v) Alloying of Gold : Pure gold is very soft and cannot be used as such for jewellery. Therefore, it is generally alloyed with other metals commonly copper or silver to make it harder and modify its colour. The purity of gold is expressed as carats. Pure gold is of 24 carat. A 18 carat gold means that is contains 18 parts of gold is 24 parts by weight of alloy. Most of the jewellery is made of 22 carat gold.

Amalgams are **homogenous mixtures of a metal and mercury.** For example, sodium amalgam contains sodium and mercury.
Different amalgams are prepared according to their used. For example,

(i) Sodium amalgam is produced to decrease the chemical reactivity of sodium metal.
It is also used as a good reducing agent.

(ii) Tin amalgam is used for silvering cheap mirrors.

(iii) The process of amalgamation is used for the extraction of metals like godl or silver from their native ores.

DAILY PRACTICE PROBLESM # 11

OBJECTIVE DPP - 11.1

1. Pure gold is equal to -
(A) 24 carat (B) 100 carat (C) 22 carat (D) 1000 carat

2. Food cans are coated with tin and not with zinc because -
(A) zinc is costlier than tin. (B) zinc has higher melting point than tin.
(C) zinc is more reactive that tin. (D) zinc is less reactive than tin.

3. Chemical rust is -
(A) hydrated ferrous oxide (B) hydrated ferric oxide.
(C) only ferric oxide. (D) None of these

4. Which of the following methods is suitable for preventing an iron vessel from rusting ?
(A) Applying grease (B) Applying paint
(C) Applying a coating of zinc (D) All the above

5. Which of the following conditions are necessary for rusting of iron?
(A) Presence of water only (B) Presence of air only
(B) Presence of water and air both (D) None of these

6. Silver metal becomes black on exposure to air by the coating of -
(A) silver chloride (B) silver oxide
(C) silver sulphide (D) silver hydroxide

7. Alloys are a homogeneous mixture of -
(A) metals only (B) non - metals only
(C) metals or metals and non-metal (D) None of these

8. German silver is an alloy of -
(A) Cu and Ni (B) Cu, Sn and Ag (C) Cu, Zn and Ni (D) Cu, Ni, Fe and Mn

9. An alloy which does not contain copper is-
(A) magnalium (B) bronze (C) brass (D) german silver

10. Which of the following is not and alloy of aluminum ?
(A) Duralumin (B) Mangalium
(C) Alnico (D) All are alloys of aluminum.

SUBJECTIVE DPP - 11.2

1. What is an amalgam ?

2. Name an alloy of copper used for making utensils.

3. Why do we make alloys ? Give two reasons.

4. Ornaments made up of gold do not get corroded. Why ?

5. Iron nails are not rusted if kept in boiled distilled water for a long time. Explain.

ANSWERS

OBJECTIVE DPP - 7.1

Ques.	1	2	3	4	5	6	7	8	9	10
Ans.	D	D	B	A	B	C	C	C	D	C

SUBJECTIVE DPP - 7.2

Sol.3 Arrangement of metals in a vertical column in the decreasing order of their chemical reactivities is called metal activity series.
Mg, Al, Zn, Fe and Cu are metals in the order of their decreasing chemical activity.

Sol. 4. (i) the copper sulphate (blue) solution gradually faded to from colourless solution.
(ii) No change takes place. It is because copper is lower is metal activity series compared to iron.

Sol.5 Gold and Silver

Sol.6 Most reactive - Na and Least reactive - Ag

Sol.7 Hydrogen gas, Fe(s) + dill. $H_2SO_4 \longrightarrow FeSO_4(aq) + H_2(g)$

OBJECTIVE DPP - 8.1

Ques.	1	2	3	4	5	6	7	8	9	10
Ans.	C	D	C	C	C	B	D	B	A	A

SUBJECTIVE DPP - 8.2

Sol.4 Because of strong electrostatic force of attraction.

Sol.5 Because of presence of free ions.

OBJECTIVE DPP - 9.1

Ques.	1	2	3	4	5	6	7	8	9	10
Ans.	D	B	C	C	B	B	B	B	B	C

SUBJECTIVE DPP - 9.2

Sol4 (i) Forth floatation process
(ii) Hydraulic washing

Sol. (i) Bauxite -$AO_2O_3.2H_2O$ (ii) Gypsum - $CaSO_2.2H_2O$
(iii) Galena - Pbs (iv) Rock salt - NaCl

OBJECTIVE DPP - 10.1

Ques.	1	2	3	4	5	6	7	8	9	10
Ans.	B	C	D	D	D	B	B	A	C	C

SUBJECTIVE DPP - 10.2

Sol.3 Zone refining and Van arkel method.

Sol.4 Distillation

Sol.5 (i) ZnO and SO_2(i) CaO and CO_2

OBJECTIVE DPP - 11.1

Ques.	1	2	3	4	5	6	7	8	9	10
Ans.	A	C	B	D	C	C	C	C	A	D

SUBJECTIVE DPP 11.2

Sol.2 Brass

Sol.4 Gold is noble metal and not affected by air and water.

Sol.5 Boiled distilled water does not contain air.